LASA ADV Chemistry
2023-2024 Workbook

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How to Use This Workbook

Welcome to Advanced Magnet Chemistry at LASA. We are excited to have you! The following guidelines will be helpful to you as you use this workbook.

- This workbook includes our practice problems and labs, but we will periodically add assignments and content as we need to. All practice problems are to be completed directly in this workbook, in your own handwriting.

- You do not have to return this workbook at all. Since it is yours forever, many students choose to reduce bulk by uncoiling them and placing sections in a binder.

- We are unable to replace paper workbooks. If you lose them, please use the electronic copy available in your teacher's BLEND course to print new paper copies from home or the library.

- We write, edit, and revise these workbooks. We have done our very best to produce a high-quality product. However, we are human and *there will be errors*. If you find errors, please notify your teacher.

- Completing the practice problems is one part of being a successful student. However, assessments require you to make connections outside an ordinary practice problem. Good scientists are good because they apply what they have learned to new situations. While all content we cover will be tested, *practice problems do not imply anything about the scope, depth or difficulty of exam or quiz questions.*

- The progression and depth of our lessons is *highly* subject to change. Please be flexible.

- We have included blank space so you can make your own notes. *Research has shown that student notetaking is an effective way to learn content, deepen understanding, and make connections to other content*.

- Please feel free to give your teacher constructive feedback on this workbook throughout the course. We hope you will enjoy chemistry!

PERIODIC TABLE OF THE ELEMENTS

Atomic number — 14
Symbol — **Si**
Atomic mass — 28.085
Name — Silicon

	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H** 1.008 Hydrogen																	2 **He** 4.0026 Helium
2	3 **Li** 6.94 Lithium	4 **Be** 9.0122 Beryllium											5 **B** 10.81 Boron	6 **C** 12.011 Carbon	7 **N** 14.007 Nitrogen	8 **O** 15.999 Oxygen	9 **F** 18.998 Fluorine	10 **Ne** 20.180 Neon
3	11 **Na** 22.990 Sodium	12 **Mg** 24.305 Magnesium											13 **Al** 26.982 Aluminum	14 **Si** 28.085 Silicon	15 **P** 30.974 Phosphorus	16 **S** 32.06 Sulfur	17 **Cl** 35.45 Chlorine	18 **Ar** 39.948 Argon
4	19 **K** 39.098 Potassium	20 **Ca** 40.078 Calcium	21 **Sc** 44.956 Scandium	22 **Ti** 47.867 Titanium	23 **V** 50.942 Vanadium	24 **Cr** 51.996 Chromium	25 **Mn** 54.938 Manganese	26 **Fe** 55.845 Iron	27 **Co** 58.933 Cobalt	28 **Ni** 58.693 Nickel	29 **Cu** 63.546 Copper	30 **Zn** 65.38 Zinc	31 **Ga** 69.723 Gallium	32 **Ge** 72.630 Germanium	33 **As** 74.922 Arsenic	34 **Se** 78.971 Selenium	35 **Br** 79.904 Bromine	36 **Kr** 83.798 Krypton
5	37 **Rb** 85.468 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.906 Yttrium	40 **Zr** 91.224 Zirconium	41 **Nb** 92.906 Niobium	42 **Mo** 95.95 Molybdenum	43 **Tc** Technetium	44 **Ru** 101.07 Ruthenium	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.76 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.90 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	71 **Lu** 174.97 Lutetium	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.84 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.23 Osmium	77 **Ir** 192.22 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** Polonium	85 **At** Astatine	86 **Rn** Radon
7	87 **Fr** Francium	88 **Ra** Radium	103 **Lr** Lawrencium	104 **Rf** Rutherfordium	105 **Db** Dubnium	106 **Sg** Seaborgium	107 **Bh** Bohrium	108 **Hs** Hassium	109 **Mt** Meitnerium	110 **Ds** Darmstadtium	111 **Rg** Roentgenium	112 **Cn** Copernicium	113 **Nh** Nihonium	114 **Fl** Flerovium	115 **Mc** Moscovium	116 **Lv** Livermorium	117 **Ts** Tennessine	118 **Og** Oganesson

Atomic masses are not listed for elements with no stable or common isotopes.

Lanthanide Series	57 **La** 138.91 Lanthanum	58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** Promethium	62 **Sm** 150.36 Samarium	63 **Eu** 151.96 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.93 Terbium	66 **Dy** 162.50 Dysprosium	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.05 Ytterbium
Actinide Series	89 **Ac** Actinium	90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactinium	92 **U** 238.03 Uranium	93 **Np** Neptunium	94 **Pu** Plutonium	95 **Am** Americium	96 **Cm** Curium	97 **Bk** Berkelium	98 **Cf** Californium	99 **Es** Einsteinium	100 **Fm** Fermium	101 **Md** Mendelevium	102 **No** Nobelium

Updated 2017

Source: International Union of Pure and Applied Chemistry

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Unit 1: Matter and Measurement

1.01: First Day Activities

Lesson 1.01 Student Notes:

1.02: Lab Safety / SDS and NFPA / “Green” Chemistry

1. (1.A) Draw a sketch of your chemistry classroom below. Design a legend. Mark the locations of these items.
    - Exits
    - Fire extinguisher
    - Safety shower
    - Fume hood
    - Fire blanket
    - Eye wash
    - Lab aprons
    - Goggles
    - Student desks

Legend:

2. (1.A) In the event of a fire, chemical spill, or other emergency, appropriate safety equipment should be easily accessible, and an emergency plan followed. What types of fire fighting equipment should be used when confronted with a chemical fire on the lab bench versus a fire on the sleeve of a classmate's clothes and write out a response plan you should follow for each scenario.

3. (1.A) Broken glassware poses a serious threat to your laboratory safety. What precautions should you take if you encounter this threat?

4. (1.A) Unlabeled chemicals are a very dangerous hazard in the lab. If you find unlabeled chemicals in your lab, how should you treat them?

Answer the following questions using the SDS (safety data sheet) for sodium hydroxide (NaOH) found at https://www.sigmaaldrich.com/US/en/sds/sigald/s8045.[1]

5. (1.B) List the chemical formula of sodium hydroxide, and other commonly used names for sodium hydroxide.

6. (1.B) What is the procedure If sodium hydroxide contacts your eyes?

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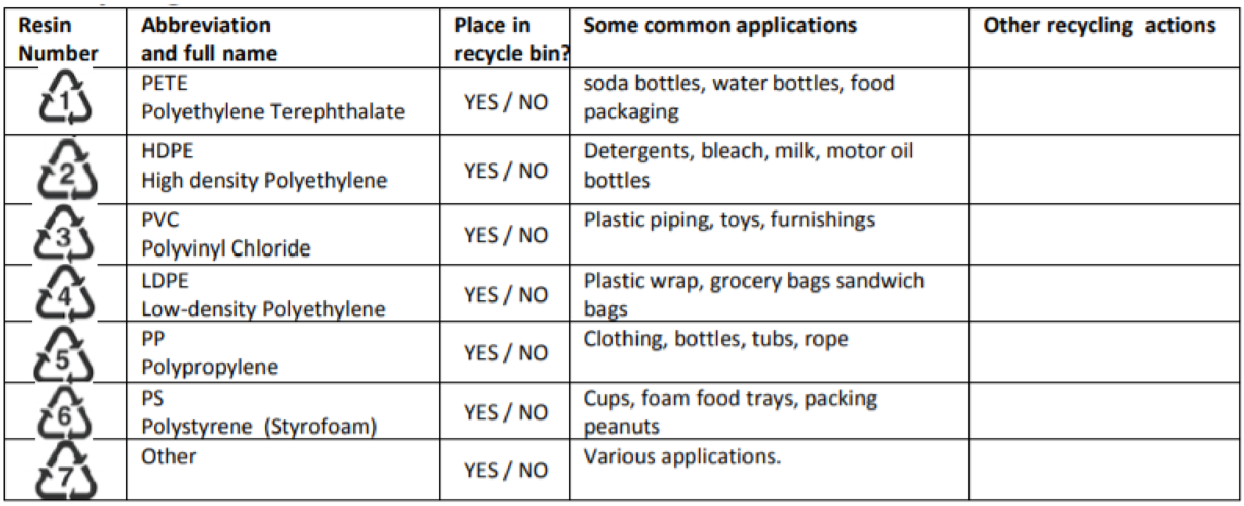
[1] "Sodium Hydroxide." *SDS Search*, Sigma-Aldrich https://www.sigmaaldrich.com/US/en/sds/sigald/s8045.

7. (1.B) What is the procedure if sodium hydroxide is swallowed?

8. (1.B) What are the NFPA ratings for Health? Fire? Reactivity? Specific hazard?

9. (1.B) What is the procedure when a contact lens wearer must work with this chemical?

10. (1.C) Using the table below, circle whether the plastic resin can be placed in a recycle bin, if you cannot place that particular plastic resin in the bin complete the “other recycling actions” columns with how and where you can recycle those plastic materials.



Resin Number	Abbreviation and full name	Place in recycle bin?	Some common applications	Other recycling actions
1	PETE Polyethylene Terephthalate	YES / NO	soda bottles, water bottles, food packaging	
2	HDPE High density Polyethylene	YES / NO	Detergents, bleach, milk, motor oil bottles	
3	PVC Polyvinyl Chloride	YES / NO	Plastic piping, toys, furnishings	
4	LDPE Low-density Polyethylene	YES / NO	Plastic wrap, grocery bags sandwich bags	
5	PP Polypropylene	YES / NO	Clothing, bottles, tubs, rope	
6	PS Polystyrene (Styrofoam)	YES / NO	Cups, foam food trays, packing peanuts	
7	Other	YES / NO	Various applications.	

11. (1.C) Describe a minimum of three ways you can reduce the amount of plastic you use every day.

12. (1.C) Describe a minimum of three ways you can reduce the amount of waste from home that makes it to a landfill.

Challenge Problems:

13. (1.A, 1.B, 1.C). Every year the chemistry teachers must complete a chemical inventory of the stockroom to record what is on the shelves. What types of chemical storage guidelines do the teachers have to follow? How do they know how to place the chemicals on the shelves? What types of safety precautions do they have to follow when placing the chemicals on shelves?

14. (1.A) There was only mention of a type ABC fire extinguisher in this safety lesson. What are all of the different types of fire extinguishers? What type of fire extinguisher would you use to extinguish a fire created by paper? What if the fire was created by chemicals?

Lesson 1.02 Student Notes:

1.03: Classification of Matter / Solids, Liquids, and Gases / Separation of Mixtures Lab

1. (4.C, 4.D) Indicate if each chemical symbol represents an element or a compound, and its state of matter. If the pure substance is a compound, state what elements are in the compound. If the pure substance is a compound, state how many atoms of each element are represented in one formula unit or molecule of that substance. (There are examples in the text.)

	$C_{(s)}$	$Br_{2\,(l)}$	$Cr_2O_{3\,(s)}$	$C_2H_{4\,(g)}$	$Sr(NO_2)_{2\,(l)}$
Element or compound?					
State of matter					
For compounds, which elements are in the compound?					
For compounds, one formula unit or molecule of this pure substance contains...					

2. (4.D) Draw a particle diagram of one formula unit/molecule of each pure substance from #1. Use any symbols and color scheme of your choice. Your particles should represent the substance; for instance, a NaCl formula unit should have one Na atom and one Cl atom stuck together and the Na and Cl symbols should look different from each other.

$C_{(s)}$ $Br_{2\,(l)}$ $Cr_2O_{3\,(s)}$ $C_2H_{4\,(g)}$ $Sr(NO_2)_{2\,(l)}$

3. (4.C, 4.D) Draw appropriate particle diagrams for samples of the following types of matter. Use any symbols and color scheme of your choice. Your particles should represent the substance; for instance, a NaCl formula unit should have one Na atom and one Cl atom stuck together and the Na and Cl should look different from each other.

A solid sample of the compound, $AgCl_{(s)}$	A gaseous sample of the diatomic element, $N_{2\,(g)}$	A liquid sample of the compound, $NaNO_{2\,(l)}$

4. (4.C, 4.D) Draw appropriate particle diagrams for samples of the following types of matter. Use any symbols and color scheme of your choice. Your particles should represent the substance; for instance, a NaCl formula unit should have one Na atom and one Cl atom stuck together and the Na and Cl should look different from each other.

A gaseous solution of two different diatomic elements	A gaseous suspension of a diatomic element and two different compounds	A liquid solution of a monatomic element, a diatomic element, and a compound

5. (4.D) How would you separate the following mixtures? Explain what you would do, with steps, in logical order; don't just write single words! (For example: First I would... then I would...)

   a) a liquid mixture of crude oil (contains gasoline, kerosene, benzene, etc. that are all liquid)

   b) a suspension of sand in sugar water (into solid sand, solid sugar, liquid water)

6. (4.C, 4.D) Classify each sample of matter thoroughly. If a question doesn't apply, leave that space blank. $SiO_{2\ (s)}$ is sand, $C_{12}H_{22}O_{11\ (s)}$ is sugar, and $Au_{\ (s)}$ and $Ni_{(s)}$ is an alloy. *Watch the state signifiers!*

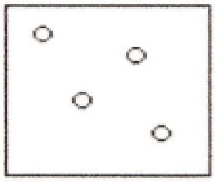
	$F_{2\ (g)}$	$Fe_2(SO_4)_{3\ (s)}$	$C_{12}H_{22}O_{11\ (aq)}$	$Au_{\ (s)}$ and $Ni_{(s)}$ in jewelry	$SiO_{2\ (s)}$ in $H_2O_{\ (l)}$
Pure substance or mixture?					
If it's a pure substance, is it an element or a compound?					
If it's an element, is it monatomic or diatomic?					
If it's a mixture, is it a homogeneous or heterogeneous?					

7. (4.C, 4.D) Classify each sample, A – P. Size differences indicate atoms of different elements. For example, sample P is: I) a homogeneous mixture, II) consists of elements, and III) a solid.

Each diagram (A - P) show a sample of substances as viewed at the atomic level. Characterize the contents of the container in terms of each of the following categories:

Category I.	Homogeneous mixture, heterogeneous mixture or pure substance
Category II.	Element(s), compound(s) or both
Category III.	Solid, liquid, gas or combination of phases

A



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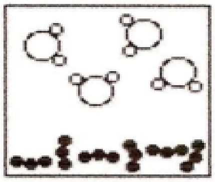

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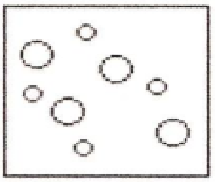

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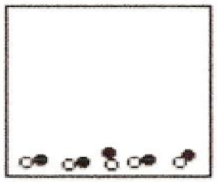

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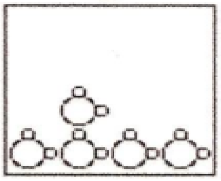

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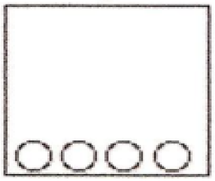

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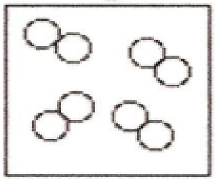

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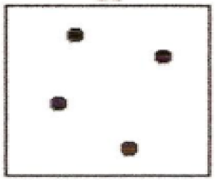

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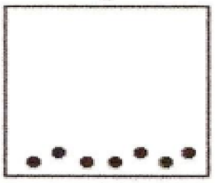

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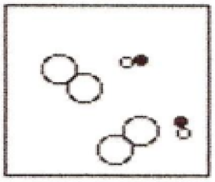

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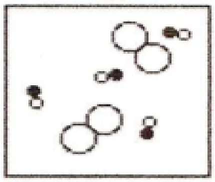

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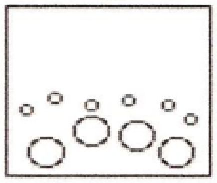

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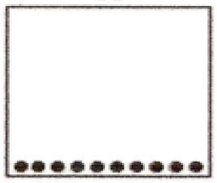

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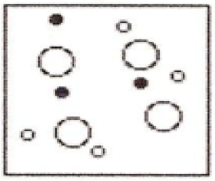

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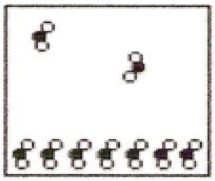

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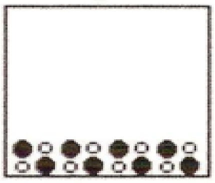


Challenge Problems:

8. (4.C) Which characteristics of a solid, a liquid, and a gas are exhibited by each of the following substances? How would you classify each substance?

    a) A bowl of pudding
    b) A bucketful of sand
    c) A root beer float

9. (4.D) Suppose a teaspoon of calcium filings and a teaspoon of powdered selenium are placed together in a metal beaker. Would this be classified as a mixture or a pure substance? Explain using complete sentences. Suppose the calcium filings and selenium are heated so that they react with each other, forming calcium selenide. Would this still be a "mixture"? Why or why not? Explain using complete sentences.

Lesson 1.03 Student Notes:

Separation of Mixtures Lab

Problem: Separate the given mixture into its components. You and your group need to present your teacher with *dry*, *solid* samples of each component of the mixture.

Materials:

Mixture
MSDS of each component of the mixture
Tap water
Beakers, funnel, filter paper, stirring rod
Magnet

Procedure: Write an organized, logical stepwise procedure that will successfully separate the components of the mixture. Hint: use as little water as possible!

Assign a role to each group member. Perform your procedure, using appropriate lab safety rules. When your procedure is complete, clean up and tidy the table.

Analysis: Your teacher will provide a sheet of paper for your group. Write the names of all group members on this sheet. Place the dry, solid samples of each component on their own circle. Label each circle with the name of the component. Have your teacher sign the paper for a clean lab table.

1.04: Physical and Chemical Properties and Changes; Phase Changes and Diagrams

(4.B) Select the best answer for each question about intensive and extensive properties. You can have more than one answer.

1. Which of the following is/are intensive properties of a box of raisins?

   a) calories per serving
   b) total grams
   c) the number of raisins
   d) the total calories

2. Which of the following is/are intensive properties of a sample of oxygen?

   a) molar mass of 31.998 g/mol
   b) volume of 150 mL
   c) gaseous at room temperature
   d) contains 150000 molecules

3. Which of the following is/are extensive properties of a sample of gold?

   a) density of 19.3 g/cm^3
   b) melts at 1064°C
   c) weighs 30.0 g
   d) conductivity of 4.10 x 10^7 S/m

4. Which of the following is/are extensive properties of coffee in a mug?

   a) total mg of caffeine
   b) g cream added per mL of coffee
   c) temperature
   d) percent sugar

5. (4.A) Determine if the unknown is a compound or a mixture based on the given description. Write a paragraph that supports your conclusion for each description.

   a) Unknown A is a cloudy orange liquid. After you filter Unknown A, large orange crystals have been left on the filter paper,and a clear liquid is in the beaker. After you boil the liquid, thin white crystals are left at the bottom of the beaker and the rest of the liquid has evaporated.

b) You heat Unknown B, a fine blue-green powder, in a Bunsen burner flame. You observe are a clear, colorless gas with an odor, a fine black powder, and a shiny brown metal.

6. (4.A) State if each equation represents a physical (phase) change or a chemical change. If it is a physical (phase) change, identify the change using the terms in the solid/liquid/gas/plasma diagram (see the textbook). (i.e. melting... vaporization... etc.)

a) $CO_{2\,(g)} \rightarrow CO_{2\,(s)}$

d) $Br_{2\,(l)} \rightarrow Br_{2\,(g)}$

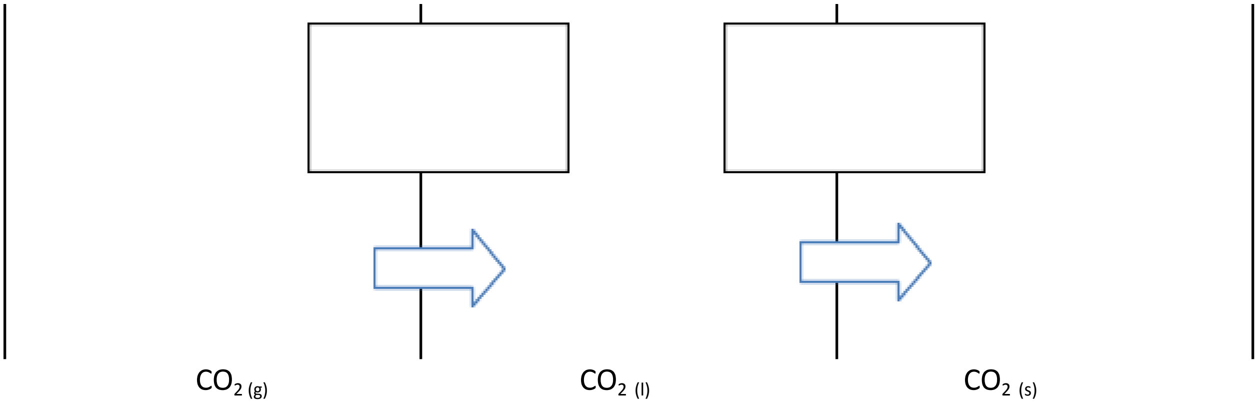
b) $CO_{2\,(g)} \rightarrow CO_{(g)} + O_{2\,(g)}$

e) $C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$

c) $I_{2\,(s)} \rightarrow I_{2\,(g)}$

f) $C_6H_{6\,(l)} \rightarrow C_6H_{6\,(s)}$

7. (4.A) Draw particle diagrams that show what happens to $CO_{2\,(g)}$ as it cools and undergoes phase changes from vapor to liquid to a solid. Indicate the changes in energy as these changes take place. See the particle diagrams in the textbook.



8. (4.A) Identify the following as being true or false to the left of the sentence.

______ a) A change in size or shape is a physical change.

______ b) A chemical change means a new substance with new properties was formed.

______ c) An example of a chemical change is when water freezes.

______ d) An example of a physical change is when platinum metal is heated, then cooled.

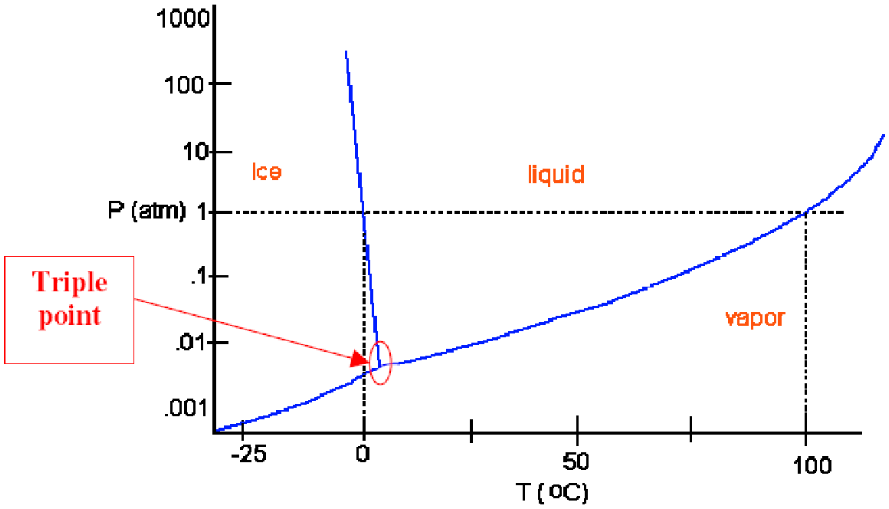
______ e) Milk turning sour is a physical change.

______ f) When citric acid and baking soda mix, carbon dioxide is produced and the temperature decreases. This must be a chemical change.

9. (4.A) State if each of these changes is a physical change or a chemical change.

a) Melting candle wax
b) Subliming iodine
c) Tarnishing silver
d) Warming water
e) Mixing blue and red colored water
f) Oxidizing copper
g) Condensing water vapor
h) Burning magnesium
i) Decomposing food
j) Changing leaf color in the fall
k) Photosynthesizing
l) Digesting food in your body

10. (3.B, 4.A) Interpret the following phase diagram (for water) by answering the questions that follow.



a) What is the freezing point and boiling point when the pressure on the water is 1 atm?

b) How does lowering the pressure affect the boiling point?

c) Estimate the boiling point of water when the pressure on the water is 0.1 atm.
d) Estimate the triple point.

e) Estimate the critical point.

f) Dry ice sublimes: that is, it changes directly from a solid to a gas. Is it ever possible for ice (water) to sublime? If so, describe the conditions necessary for sublimation to occur.

g) Is it ever possible for solid H_2O to exist at a temperature above 0°C? Explain.

Challenge Problems:

11. (4.A, 4.D) If a piece of hard white sidewalk chalk is heated strongly in a flame, the mass of the piece of chalk will decrease, and eventually the chalk will crumble into a fine white dust. Does this change suggest that the chalk is composed of an element or a compound? Explain.

12. (4.A) If you place a glass rod over a burning candle, the glass appears to turn black. What is happening to each of the following (physical change, chemical change, both, or neither) as the candle burns? Explain each answer using complete sentences.

    a) The wax
    b) The wick
    c) The glass rod

13. (4.A) During a very cold winter, the temperature may remain below freezing for extended lengths of time. However, fallen snow can still disappear, even though it cannot melt. This is possible because a solid can vaporize directly without passing through the liquid state. What is this process called? Is this a physical or chemical change? Explain, using complete sentences.

Lesson 1.04 Student Notes:

1.05: The SI Measurement System and Scientific Notation

1. (2.G) Express the following numbers in scientific notation.

5800000000 =	0.000192 =
34.5×10^{3} =	0.004×10^{-3} =
0.72×10^{-6} =	0.029×10^{2} =

2. (2.G) Solve the following problems and express your answers in scientific notation.

$(5.4 \times 10^{-8}) + (6.6 \times 10^{-9})$ =	$(4.4 \times 10^{5}) - (6.0 \times 10^{6})$ =
$(3.24 \times 10^{4}) + (1.1 \times 10^{2})$ =	$(0.434 \times 10^{-3}) - (6.0 \times 10^{-6})$ =

3. (2.G) Solve the following problems. All answers should be left in scientific notation.

$(3.4 \times 10^{3})\ (2.0 \times 10^{7})$ = ______________	$(5.4 \times 10^{2})\ /\ (2.7 \times 10^{4})$ = ______________
$(4.6 \times 10^{1})\ (6.7 \times 10^{4})$ = ______________	$(8.4 \times 10^{-3})\ /\ (4.0 \times 10^{5})$ = ______________

4. (2.G) Combine everything you have learned and perform the following calculation. Write your answer in scientific notation.

$(3.24 \times 10^{8})\ (14000)\ /\ (3.5 \times 10^{-3})$ = __________________

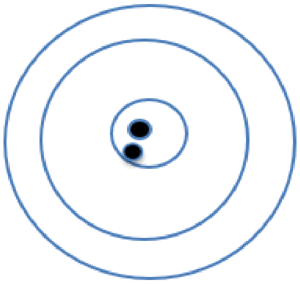
Challenge Problem:

5. (3.A) As you know, there are different measuring systems in use around the world. Research one of the alternative measurement systems. (You may not research the metric or Imperial systems). Evaluate and analyze this alternative system. How does it relate or not relate to measurement systems you have used and been taught?

Lesson 1.05 Student Notes:

1.06: Errors and Significant Figures/Accuracy and Precision of Glassware Lab

1. (2.F, 3.B) Use the dartboards to illustrate your knowledge of the concepts of accuracy and precision. One dot represents a dart. One is done for you as an example.

	
Accurate and precise	Not accurate and not precise
Accurate and not precise	Not accurate and precise

2. (3.B) Would the following events be systematic errors, random errors, or neither?

________ a) The reading on your electronic balance fluctuates when air currents move by.
________ b) You forgot to tare your electronic balance between measurements.
________ c) You measure a volume correctly but record the measurement incorrectly.
________ d) You do not measure volume or length using one uncertain digit.
________ e) You calculate a density value incorrectly.

3. (2.G, 3.B) You measure the boiling point of ethyl alcohol to be 77.0°C, but the actual boiling point is 80.0°C. What is the percent error?

4. (2.G, 3.B) An object has a mass of 35.00 grams. On your balance, it weighs 34.85 grams. Calculate the percent error of your measurement.

5. (2.G, 3.B) The density of water at 4°C is known to be 1.00 g/mL. If the experimentally found density of water is 1.08 g/mL, what is your percent error?

6. (2.G, 3.B) The uncertainty of a balance is ± 0.01 g. Write the numbers that should be recorded as data with the correct number of significant figures. Some values may already be correct.

445.81 g ________________ 6.731 g ________________ 5.555 g ____________________

5872.30 g _______________ 5.556 g ________________ 5.505 g ____________________

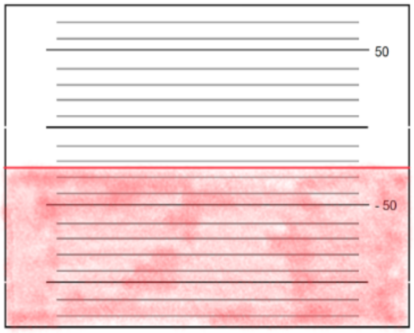
7. (2.G, 3.B) Determine the correct number of significant figures in the following numbers.

10.01 g __________________________ 140 g _________________________

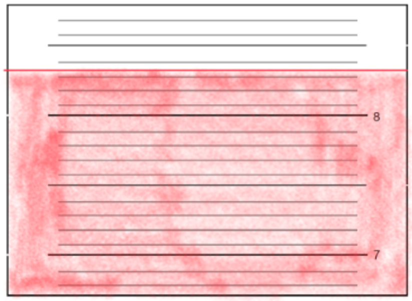
0.0010 g _________________________ 140.0 g ________________________

1.100 g __________________________ 1100 g _________________________

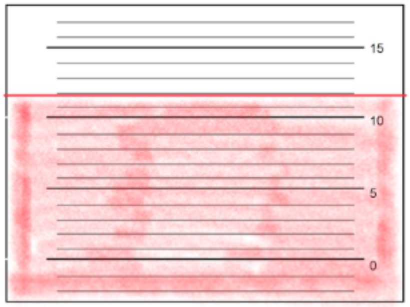
8. (2.F, 2.G, 3.B) Label the temperature measurement from each thermometer below.



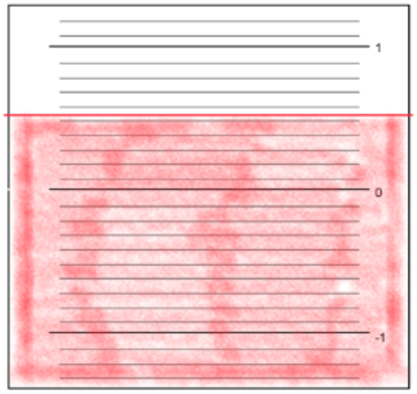
a) ________________________



b) ________________________

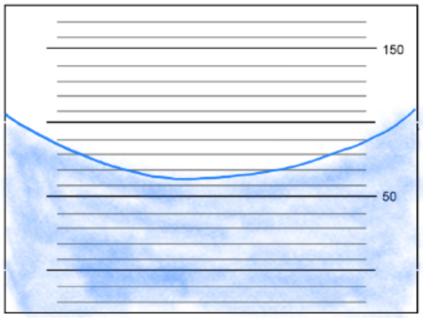


c) ________________________

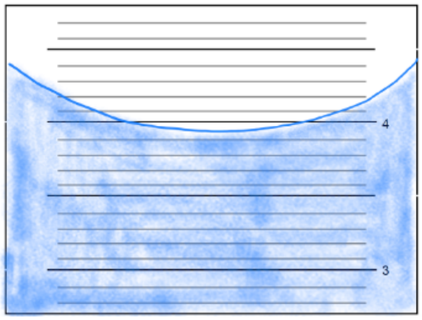


d) ________________________

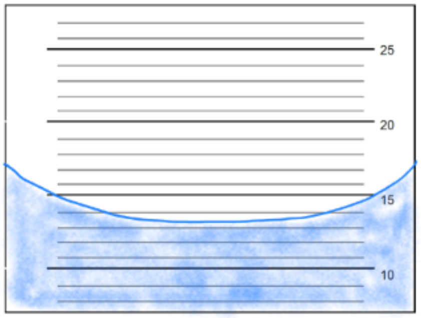
9. The graduated cylinders below have different precisions. Label the amount of water in each of the four graduated cylinders to the correct number of significant figures.



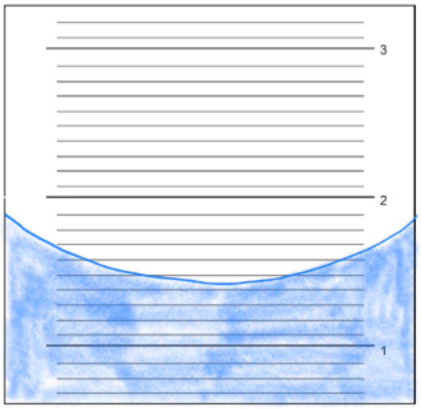
a) ______________________



b) ______________________



c) ______________________



d) ______________________

10. (2.G, 3.B) Solve the following problems and report your answer with the correct number of significant figures and units. Round if necessary.

    16.0 g + 3.106 g + 0.8 g =

    9.002 m^2 - 3.10 m^2 =

    39.64 L + 1.3 L – 25.394 L =

11. (2.G, 3.B) Solve the following and report your answer with the correct number of significant figures and units.

4.5 m × 6.75 m =

3.0 cmol / 230.5 cmol =

0.0847 g × 31000 g =

12. (2.G, 3.B) Solve the following and report your answer with the correct number of significant figures.

(6.8 + 0.047) × 17.44 =

(320. - 22.7) / 3.8 =

(14.86 + 13.7) × (65.346 - 4.10) / (0.07 × 32.888) =

Note: From this point forward until the end of the course, all numerical answers MUST be expressed and rounded using the appropriate significant figure rules.

Challenge Problems:

13. A 25.00 g sample of a solid is placed in a graduated cylinder, and then the cylinder is filled to the 50.0 mL mark with benzene (a liquid). The mass of benzene and solid together is 58.80 g. If the solid does not dissolve in benzene and the density of benzene is 0.880 g/cm^3, calculate the density of the solid.

14. You have two beakers, one filled to the 100 mL mark with sugar (the sugar has a mass of 180.0 g) and the other filled to the 100 mL mark with water (the water has a mass of 100.0 g). You pour all the sugar and all the water together in a bigger beaker and stir until the sugar is completely dissolved.

    a) Which of the following is true about the mass of the solution? Explain using complete sentences.
        i) It is much greater than 280.0 g.
        ii) It is somewhat greater than 280.0 g.
        iii) It is exactly 280.0 g.
        iv) It is somewhat less than 280.0 g.
        v) It is much less than 280.0 g.

    b) Which of the following is true about the volume of the solution? Explain using complete sentences.
        i) It is much greater than 280.0 mL.
        ii) It is somewhat greater than 280.0 mL.
        iii) It is exactly 280.0 mL.
        iv) It is somewhat less than 280.0 mL.
        v) It is much less than 280.0 mL.

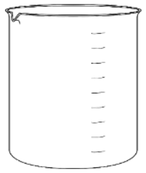
15. The density of an irregularly shaped object was determined as follows: the mass was found to be 28.90 g ± 0.03 g. A graduated cylinder was partially filled with water. The reading of the level of the water was 6.4 cm^3 ± 0.1 cm^3. The object was dropped in the cylinder, and the level of the water rose to 9.8 cm^3. What is the density of the object with appropriate error limits?

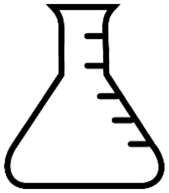
Lesson 1.06 Student Notes:

Accuracy and Precision of Glassware Lab

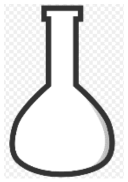
Problem: Which piece of equipment – the 50 mL beaker, 50 mL volumetric flask, the 25 mL graduated cylinder or the 50 mL Erlenmeyer flask – is the most accurate and precise at measuring volume?

Materials:
50 mL beaker
50 mL volumetric flask
Electronic balance (use the SAME balance every time)
50 mL Erlenmeyer flask
25 mL graduated cylinder
Paper towels
Water


Beaker


Erlenmeyer flask

Graduated cylinder


Volumetric flask

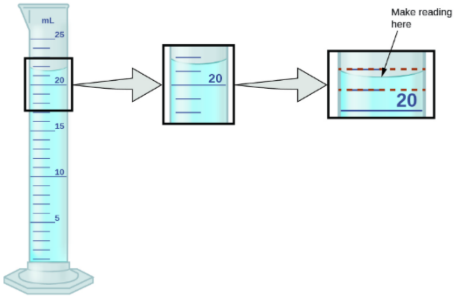
Procedure:

Split your lab group in half. Each lab group half will take two (2) of the glassware options. Repeat steps #1-5 for each piece of glassware.

1. Place a dry, empty 50 mL beaker on an electronic balance. **Tare** the balance. This means to "zero it out"; press Zero or Tare to reset it to 0. Record its mass.

2. Use the lines on the beaker to measure 20 mL of water directly into the beaker. DO NOT USE ANY OTHER GLASSWARE OTHER THAN THE BEAKER. Read the volume at the bottom of the liquid level, or the **meniscus**.

   *Always read volumes with the number of digits that the instrument gives you *and then estimate one more*. For example, if the beaker meniscus is just above "10 mL" and the lines go 10, 20, 30, etc. record the volume as "11 mL" or "12 mL" (depending on your judgement). On the graduated cylinder, if you see the liquid meniscus is on the 11 mL line, you record the volume as 11.0 mL. This last digit is called an **uncertain digit** and is included in your recording.



[2]**To determine the number of digits that would be recorded in the measurement of the graduated cylinder to the left, first determine where the meniscus is, then determine what the certain digits are by determining what measurement the markings occur, in this case the markings are at every 1 mL. You can be certain that the meniscus is between 21 mL and 22 mL; therefore, the certain digits are 21 mL. The uncertain digit is the estimated digit (depends on the eyes of the person reading the volume). In this case, the estimation is between 0.5 mL and 0.9 mL, let's say 0.6 mL. The final answer would include the two certain digits and the one uncertain digit, 21.6 mL.

3. Weigh the beaker with the water inside it. Record the mass of the beaker plus the water.

4. Pour the water out into the sink. (Yes, really.) Dry the beaker well.

5. Calculate the mass of the "20 mL" portion of water you measured.

6. Repeat steps 1-5 for the other three pieces of glassware.

7. The density of water is 1 g/mL; that is, 1 g of water has a volume of 1 mL. Calculate the *real* volume of the "20 mL water" for each of the pieces of glassware since you have the mass of each portion of water.

Data

Accuracy and Precision of a 50 mL Beaker	
Mass of empty 50 mL beaker (g):	
Mass of dry beaker + "20 mL" water	
Mass of "20 mL" water	
Real volume of "20 mL" water	

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[2] OpenStax. "1.5 Measurement Uncertainty, Accuracy, and Precision." *Chemistry*, Rice University, 2016, opentextbc.ca/chemistry/chapter/measurement-uncertainty-accuracy-and-precision/. Licensed under CC BY 4.0: https://creativecommons.org/licenses/by/4.0/legalcode.

Accuracy and Precision of a 50 mL Volumetric Flask	
Mass of empty 50 mL volumetric flask (g):	
Mass of dry volumetric flask + “20 mL” water	
Mass of “20 mL” water	
Real volume of “20 mL” water	

Accuracy and Precision of a 25 mL Graduated Cylinder	
Mass of empty 25 mL graduated cylinder(g):	
Mass of dry graduated cylinder + “20 mL” water	
Mass of “20 mL” water	
Real volume of “20 mL” water	

Accuracy and Precision of a 50 mL Erlenmayer Flask	
Mass of empty 50 mL Erlenmayer flask (g):	
Mass of dry Erlenmayer flask + “20 mL” water	
Mass of “20 mL” water	
Real volume of “20 mL” water	

Conclusion: Which is the most accurate and precise glassware of the four pieces that you selected? Write a conclusion in paragraph form. Clean up and dry all glassware using the white basket items at the table and put everything back into the bin.

1.07: Dimensional Analysis

Solve all problems using dimensional analysis. You will need to know and memorize the metric prefixes and conversion factors in your electronic version of your course notes.

1. (2.G, 3.B) Convert 0.080 µm to mm.

2. (2.G, 3.B) Convert 350 mm to km.

3. (2.G, 3.B) Convert 7.50 x 10^{-2} kg to µg.

4. (2.G, 3.B) Convert 0.001705 nL to cL.

Solve all problems using dimensional analysis. You need to know and memorize metric prefixes and conversion factors. Also, be aware that density is a conversion factor! **(You ARE expected to know and use the density of water, 1 g/mL, if you need to.)**

5. (2.G, 3.B) Convert 555,000. square centimeters to square miles.

6. (2.G, 3.B) Convert 15800000 cubic millimeters to cubic meters.

7. (2.G, 3.B) Convert 30.0 m/s to km/hr.

8. (2.G, 3.B) Convert 459 ft/sec to mi/hr.

Challenge Problems:

For the questions that follow, use the conversion factors in **9.** and/or **10.** to aid in dimensional analysis. You need to use the metric prefixes and conversion factors given in your course notes.

9. (2.G, 3.B) 1 mole of gold contains 6.02×10^{23} atoms. How many moles is 2.00×10^{22} atoms of gold?

10. (2.G, 3.B) 1 mole of gold weighs 196.966 g. How many atoms are in 800 mg of gold?

11. (2.G, 3.B) What is the mass of 3.75 x 10^{24} gold atoms in cg?

12. (2.G, 3.B) The density of mercury (Hg), a monatomic element, is 13.6 g/mL. 1 mole of mercury has a mass of 200.59 g and has 6.02 × 10^{23} mercury atoms. You have a sample of mercury that is 75% pure. How many atoms are in 50. mL of your mercury sample?

13. (2.G, 3.B) The density of a 6% solution of sulfuric acid, dissolved in water, is 1.04 g/mL. ("6%" means that 6% of the solution consists of pure sulfuric acid.)

    a) What is the mass of 500. mL of a 6% solution of sulfuric acid?

    b) How many milliliters of the 500. mL sample consist of pure sulfuric acid?

    c) Given that 1 mole of pure sulfuric acid weighs 98.078 g, calculate the mass of the pure sulfuric acid in the 500. mL sample.

Lesson 1.07 Student Notes:

Unit 2: Atomic Structure and the Periodic Table

2.01: Early Atomic History

1. (2.B) Could Aristotle and Democritus' ideas be classified as scientific hypotheses? Justify your answer in a well-written paragraph.

2. (2.B, 2.C, 2.D) Could you classify the law of conservation of mass and the law of definite proportions as scientific hypotheses, scientific theories, or neither? Justify your answer with a well-written paragraph, clearly showing your understanding of how scientific hypotheses and scientific theories differ.

3. (2.C, 2.D) Explain why Dalton's atomic theory is a scientific theory, not a scientific hypothesis.

4. (3.F, 6.A) Dalton's model of the atom could be visualized as a hard, round sphere. Is this similar or different to Democritus' model, and how? Explain your answer.

5. (3.F, 6.A) Two of Dalton's postulates come directly from the work of other scientists explored in this lesson. Explain the work and conclusions of these other scientists and how they support two of Dalton's postulates, in a well-written paragraph.

6. (2.H, 2.I, 3.A, 6.A) Draw Thomson's cathode ray tube experimental results If the cathode ray tube particles were *positively* charged.

7. (3.F, 6.A) Explain how Thomson’s, Rutherford’s and Chadwick’s experimental and observational testing and conclusions, as described in this lesson, have added to Dalton’s atomic model.

Thomson:

Rutherford:

Chadwick:

8. (2.H, 3.A, 6.A) Some of Dalton’s Postulates are now incorrect as a result of Thomson, Rutherford, and Chadwick’s conclusions. Briefly describe those postulates and explain how these conclusions contradict them, in a well-written paragraph.

9. (2.H, 2.I, 3.A, 6.A) Describe the expected results of Rutherford et. al.’s gold foil experiment if… (all else being the same)

a) they used platinum foil (atomic number = 78)

b) they used thallium foil (atomic number = 81)

10. (3.B) Fill in the table below with the correct values.

Symbol	Atomic number	Mass number	Charge	# of nucleons	# of protons	# of neutrons	# of electrons
${}^{39}_{19}K$							
${}^{41}_{19}K$							
${}^{41}_{19}K^{1+}$							
	12	25	0				
			1-	35			18
					7	7	10

Challenge Problems:

11. (3.A, 3.B, 3.F) In 1908, Robert Millikan began work on determining the exact amount of charge on an electron. The experiment he and his colleagues worked on is known as the "oil-drop experiment." Research this experiment and its conclusions. How does this experiment and its conclusions support Thomson's atomic model?

12. (3.A, 3.B, 3.F) In 1969, Murray Gell-Mann won the Nobel Prize in Physics for rethinking the composition of matter and coined the term for a brand-new subatomic particle – the "quark." Research the "quark". Is the existence of quarks supported by the experimental design and conclusions you have learned about in this section, and to what extent? Justify your answer.

Lesson 2.01 Student Notes:

2.02: Isotopes, Average Atomic Mass and Mass Spectrometry

1. (3.A, 3.F, 6.A) The existence of isotopes effectively disproves one of Dalton's Postulates. Describe this postulate and explain how the existence of isotopes disproves the postulate.

2. (2.G, 6.C) Use the data below to calculate the average atomic mass of magnesium in amu.

Isotope name	Percent abundance	Isotope atomic mass (amu)
Mg-24	78.99%	23.985
Mg-25	10.00%	24.986
Mg-26	11.01%	25.983

3. (2. G, 6.C) Calculate the average atomic mass of iron in amu, using information in the table.

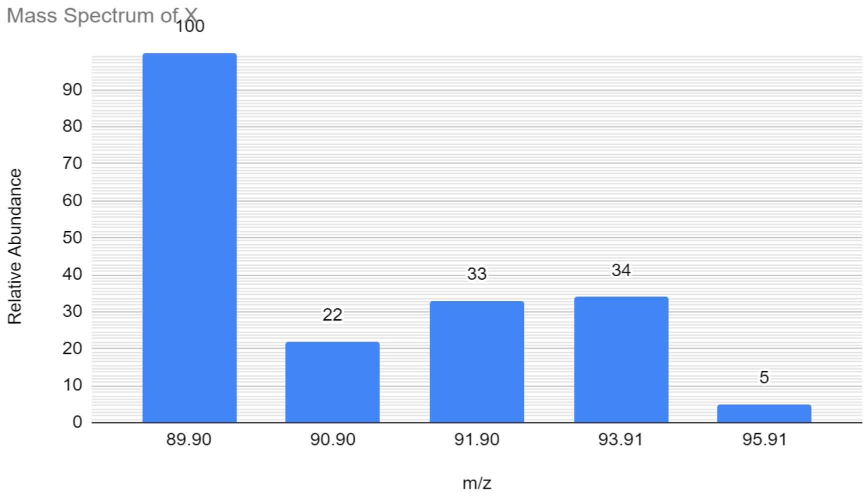
Isotope name	Percent abundance	Isotope atomic mass (amu)
Iron-54	5.90%	53.94
Iron-56	91.72%	55.93
Iron-57	2.10%	56.94
Iron-58	0.280%	57.93

4. (2.G, 3.B) Gallium consists of two naturally occurring isotopes, Ga-69 and Ga-71, with respective masses of 68.926 and 70.925 amu. Calculate the percent abundance of each isotope.

5. (2.G, 3.B) Copper is made up of two isotopes, Cu-63 (62.9296 amu) and Cu-65 (64.9278 amu). Given copper's atomic weight of 63.546, calculate the percent abundance of each isotope.

6. (2.G, 3.B) Strontium consists of three isotopes with atomic masses of 86, 87, and 88 amu respectively. Sr-88 is 72.6% abundant. If the average atomic mass of strontium is 87.62 amu, what are the percent abundances of Sr-86 and Sr-87?

7. (2.G, 2.H, 6.C) Consider the mass spectrum for element X.



a) Calculate the average atomic mass of this element.

b) In a well-written paragraph, summarize the properties of Element X in terms of: the number of isotopes Element X has, which isotopes contribute the most and the least to the average atomic mass, and the identity of element X.

8. (2.G, 2.H, 6.C) Lead has four naturally occurring stable isotopes. The table shows their masses and relative abundances.

Mass number	Relative abundance (%)
204	1.4
206	25.0
207	21.6
208	52.0

a) In the space below, draw a mass spectrum for lead, based on the data in the table. (You don’t *have* to have one of the abundances be 100%.)

b) In the space below, calculate the average atomic mass of lead in amu.

Pennium Lab

Problem*:* Plan and implement an investigative procedure to collect data and calculate, WITHOUT OPENING THE CONTAINER, the number of pre-1982 and post-1982 pennies in your container. There are two isotopes, “pre-1982” and “post 1982”.

VERY IMPORTANT! You must use this information:

- There are 10 pennies in a container.
- The mass of one pre-1982 penny: 3.1 g
- The mass of one post-1982 penny: 2.5 g

Procedure: You have access to an electronic balance. Write a logical step by step procedure that will help you collect the data you need to solve this problem.

Data*:*

Mass of empty container + lid (written on container): _____________ g

Mass of container + 10 pennies (weigh to hundredths place): ____________ g

Calculations*:* Show all the work!

a) Calculate the mass of the 10 pennies without the container mass.

b) Write a mathematical formula to solve for the number of pre-82 pennies and the number of post-82 pennies. The “atomic mass” of each pennium isotope is written near the top of this page. (Hint: Let “x” stand for the number of one of the isotopes.)

c) Use your formula to solve for the isotopic abundance. Then, calculate the number of pre-82 pennies and post-82 pennies in your can. Round up to the nearest penny.

Challenge Problems:

9. (2.G, 3.B) The average atomic mass of one silicon atom is 28.015 amu. Use dimensional analysis to show that this value is numerically equal to the molar mass of silicon in g/mol. Use these conversion factors: 1 amu = 1.66054×10^{-24} g and 1 mole = 6.022×10^{23} atoms.

10. (2.G, 3.B) Element E has an average atomic mass of 90.0 amu. It has three isotopes with atomic masses of 89.5 amu, 89.8 amu and 90.3 amu respectively. If the first isotope has a percent abundance of 10.0%, calculate the mass of each isotope in a 50.0 g sample of naturally occurring Element E.

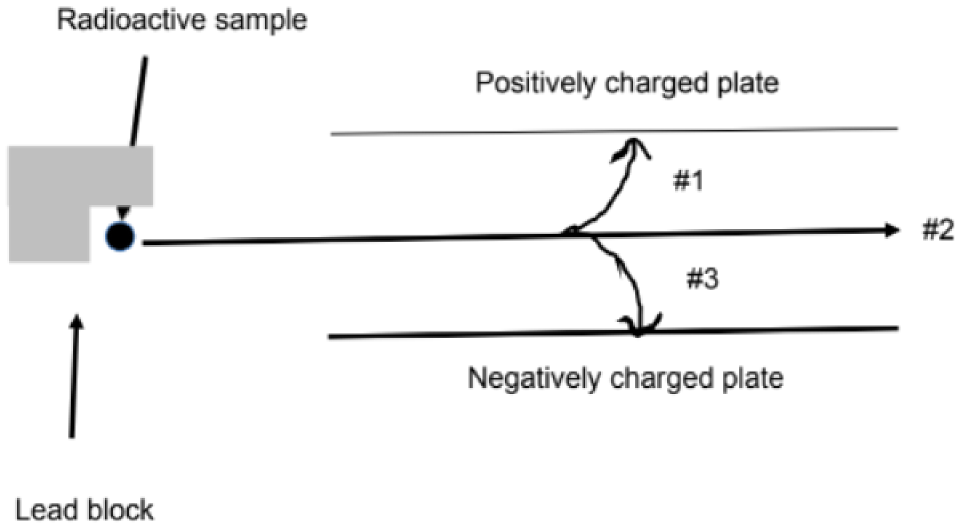
Lesson 2.02 Student Notes:

2.03: Nuclear Chemistry

1. (3.B, 12.A) Compare alpha, beta, and gamma radiation in terms of mass, charge, energy, and penetrating power.

Particle	Mass	Charge	Energy	Penetrating Power
Alpha				
Beta				
Gamma				

2. (2.H, 2.I, 12.A) What type of radioactive particle(s) could be #1, #2, or #3? Justify each answer.



3. (12.A) Complete the following nuclear equations to balance them. Identify the type(s) of nuclear decay represented by the equation.

Type(s) of Nuclear Decay

a) ${}^{235}_{92}U \rightarrow$ ________ $+ \; {}^{231}_{90}Th$

b) ${}^{142}_{61}Pm \; +$ ________ $\rightarrow {}^{142}_{60}Nd$

c) ${}^{131}_{53}I \rightarrow$ ________ $+ \; {}^{0}_{-1}e \; + \; {}^{0}_{0}\nu$

4. (12.A) Write balanced nuclear equations for each process.

a) The alpha decay of radon-198.

b) The beta (electron) decay of uranium -237.

c) Positron emission from silicon-26.

d) The electron capture of thorium-235.

5. (3.B, 12.B) Compare the processes of nuclear fusion and nuclear fission in terms of: their definitions, examples, the amount of energy involved, and whether nuclear waste is produced.

6. (12.A, 12.B) Complete the following nuclear equations to balance them. Then, identify each equation as nuclear fusion or nuclear fission.

Type of Reaction

a) $^{37}_{17}Cl + \_\_\_\_\_\_ \rightarrow ^{35}_{16}S + ^{4}_{2}He$

b) $^{14}_{7}N + ^{1}_{0}n \rightarrow \_\_\_\_\_\_\_ + ^{1}_{1}H$

c) $^{239}_{92}Pu + \_\_\_\_\_\_\_ \rightarrow ^{0}_{-1}e + ^{241}_{95}Am$

d) $^{63}_{29}Cu + ^{2}_{1}H \rightarrow \_\_\_\_\_\_\_\_ + ^{1}_{0}n$

7. (12.A, 12.B) Write balanced nuclear equations for each process.

a) The alpha particle bombardment of plutonium-239 forms three products. Two of these reaction products are a hydrogen-1 atom and 2 neutrons.

b) The neutron-induced fission of U-235 forms Sr-90 and one other product.

c) Two hydrogen-2 atoms fuse to form helium-3 and one other product.

8. (2.G) Fe-61 has a half-life of 6.00 min. Of a 100.0 mg sample, how much will remain after 18.0 min?

9. (2.G) After 20.0 days, a 120 kg sample of Bi-210 decays down to just 7.5 kg. What is its half-life?

10. (2.G) What percent of a sample of a radioactive element whose half-life of 5.0 years will decay after 25 years? (Hint: start with A_0 = 100.)

11. (2.G) How long will it take for an 80.0 g sample of cobalt-60 to decay down to 13.0 g? (half-life = 5.27 years)

12. (2.G) An ore sample is found to contain 6.78 g of K-40. How much did it contain 6.0 billion years ago? (half-life = 1.25×10^9 years)

Challenge Problems:

13. (12.A) Predict a likely mode of decay – alpha, beta, positron, or electron capture – for each of the following parent nuclides. Write and balance a nuclear equation based on the mode of decay you chose. (Hint: Calculate the neutron-to-proton ratio first.)

	Type of Decay	Decay Equation
a) Ru-90		
b) P-27		
c) Fe-45		
d) Mo-109		

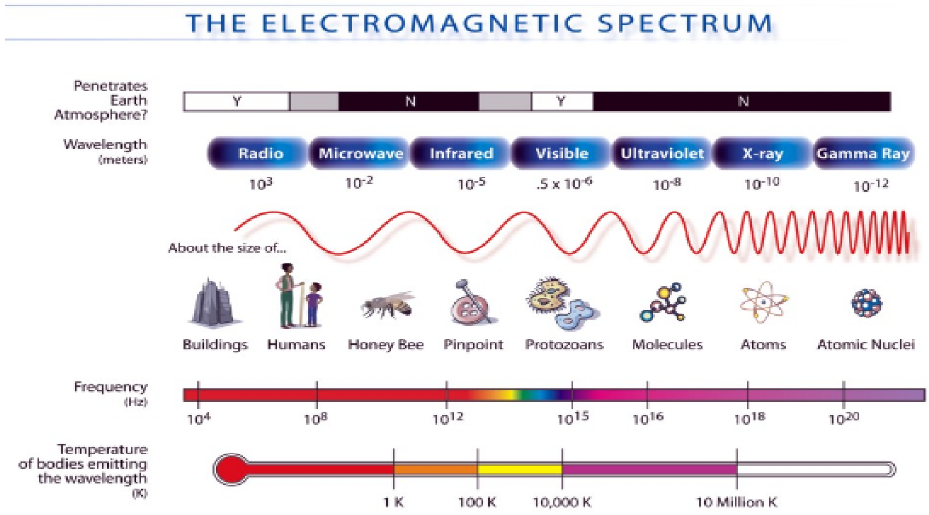
14. (3.A, 3.B, 3.F) Rutherford did extensive work on radioactivity, including the first descriptions of alpha, beta, and gamma radiation, as well as the half-life concept. How does the radioactivity concept support – or not support – Dalton's atomic theory of matter? Make reference to specific Dalton's Postulates in your answer.

15. (3.B, 3.E, 3.F) Research the uses and applications of radioactivity in medicine. Select one application and summarize your research in a well-written paragraph.

Lesson 2.03 Student Notes:

2.04: Bohr's Atomic Model

Use this graphic[3] to answer the following question:



1. (6. B) Refer to the graphic above to answer each question in a brief, well-written sentence. For example: "As _____ increases, _______ decreases, and vice versa."

a) Describe the mathematical relationship between energy and frequency.

b) Describe the mathematical relationship between energy and wavelength.

c) Describe the mathematical relationship between wavelength and frequency.

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[3] Science Mission Directorate. "Introduction to the Electromagnetic Spectrum" NASA Science. 2010. National Aeronautics and Space Administration. Retrieved June 11, 2020. http://science.nasa.gov/ems/01_intro.

2. (2.G, 6.B) Calculate the wavelength, in **m**m, of radiation with a frequency of 2.50×10^{13} Hz.

3. (2.G, 6.B) Calculate the frequency of radiation with a wavelength of 1500 nm.

4. (2.G, 6.B) Which type of radiation has a longer wavelength: red visible light (with a frequency of 4.3×10^{14} Hz) or blue visible light (with a frequency of 6.7×10^{14} Hz)? Support your answer with calculations.

5. (2.G, 6.B) A line in the atomic emission spectrum of hydrogen has a wavelength of 250. nm. Calculate the frequency of one quantum of this radiation, and the energy of this quantum in Joules.

6. (6.A) Draw a three- or four-panel “cartoon” (really, a cute particle diagram) that illustrates the process of a ground state electron becoming excited and then falling back down to the ground state, releasing quantized energy in the process.

7. (2.G, 3.B, 3.F) Calculate the wavelength AND the energy of one photon emitted as an electron falls from $n = 4$ to $n = 3$.

Wavelength calculation:

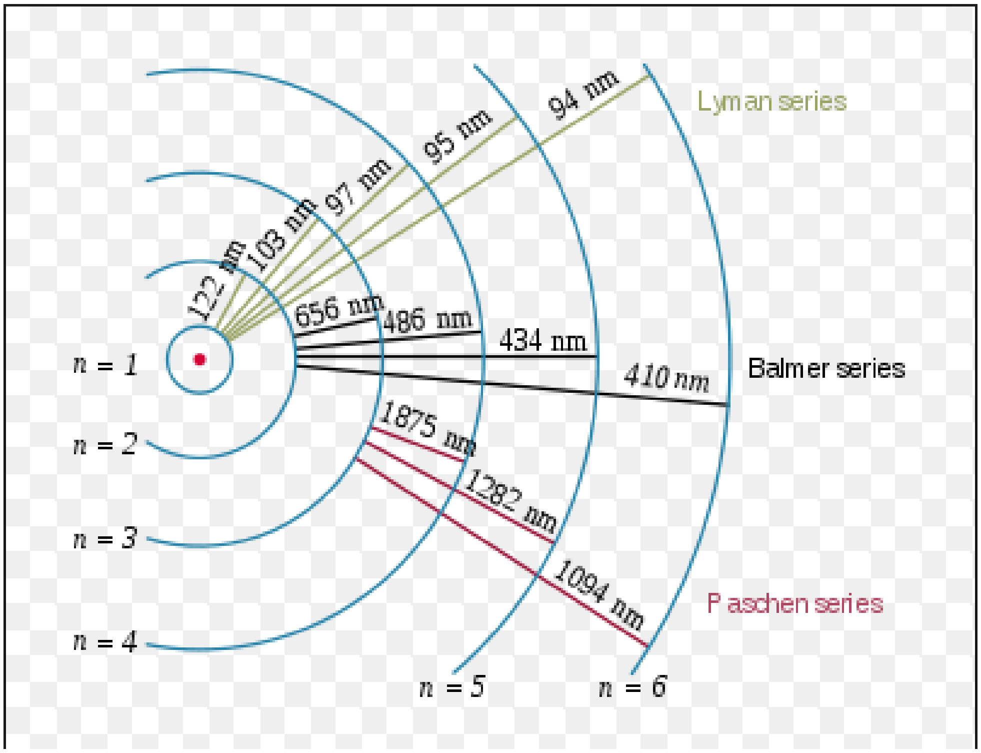
Energy calculation:

8. (3.B, 3.F, 6.A) Energy levels ($n = \ldots$) can *only* be positive whole numbers. Use this fact to explain how the Rydberg equation supports Bohr’s atomic model.

Challenge Problem:

Read the text below and answer the question that follows.

Early spectroscopists examined hydrogen spectral lines and defined “series” of lines where all the transitions ended at one energy level. For instance, the Lyman series shows the wavelengths for all the transitions to a final energy level of n = 1. The Balmer series transitions all end at *n* = 2, and the Paschen series transitions all end at *n* = 3.



“Hydrogen Spectral Series.” *Wikipedia*, Wikimedia Foundation, 28 June 2019, en.wikipedia.org/wiki/Hydrogen_spectral_series. Licensed under CC BY 2.5: https://creativecommons.org/licenses/by/2.5/legalcode

Since the Rydberg equation was able to predict the wavelengths of these transitions, and thus back up experimental data, this was used to further demonstrate the existence of energy levels, that each energy level has associated with it a discrete energy value, and that each transition has an associated wavelength associated with it. Once you get the wavelength, you can, if you need to, calculate the energy and the frequency of that quantum using the equations in this section.

9. (2.G, 3.F, 8.A) The "Balmer series" describes spectral line emissions in a hydrogen atom which correspond to "falls" by excited electrons from higher energy levels to *n* = 2. For *one* of these four "falls" (you choose), calculate the energy (in electron volts) of one mole of this radiation and the frequency of the released radiation. The wavelength of each fall is in the graphic on the previous page.

Energy in eV of one mole of this radiation:

Frequency of radiation:

Lesson 2.04 Student Notes:

2.05: Flame Tests Lab

Problem:

In this activity, you will investigate the colors of flame produced by solutions of metal salts. The energy of the emitted photons determines the color of light observed in the flame, so you must be very descriptive in your observations. By comparing the color given off by an unknown with the known metal salts, the identity of the metal salt can be determined. You will also be able to infer whether relying on the results of a flame test will definitively identify an unknown. You can even use this method to investigate the identities of metal salt mixtures.

Materials:

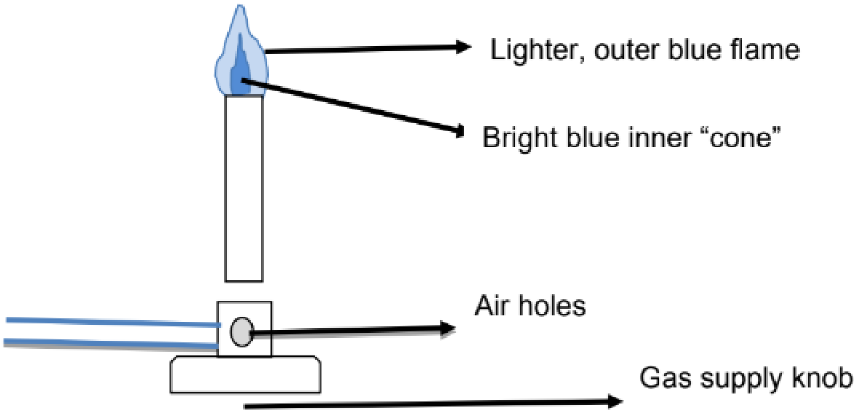
- Matches
- Bunsen burner and gas source
- Bottle(s) containing labelled metal salt solutions
- Bottle(s) containing labelled unknowns and/or mixtures (coded by letter)
- Nichrome wire loop(s)
- Appropriate safety equipment

Procedure:

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

1. At each lab station you will find one to two bottles of a metal solution and/or an unknown, identified by number (#1 through 6). The unknown may either be a single unknown or a mixture of unknowns. Please check the labels to determine the identity of the solution(s).

2. Light your Bunsen burner. It should be burning the characteristic “natural gas blue” color. If any other color is observed, adjust the air hole. You need a *bright* blue cone to appear at the center of the flame. Yellow flames are too cool. You can adjust the gas supply by turning the knob underneath the Bunsen burner.



3. If your flame blows out, TURN OFF the gas immediately and relight the burner.

4. Take a nichrome wire loop and heat the wire until it is "clean" - you should see a yellow or colorless flame if it is clean. Allow the loop to cool. Don't put a hot loop on the table or in the plastic lab tray.

5. Dip the loop into ONE of the labelled "known" bottles. DO NOT cross contaminate**.** Put the soaked wire back into the flame.

6. Record the color and intensity of the resulting flame. Be descriptive. "Red" is not descriptive. "Bright cherry red" or "pale orange-red" is descriptive.

7. Repeat steps 1-6 for any other bottles of solution at that same lab table, including any unknowns.

8. Move to the next station(s) and repeat steps #1-6 for the new solutions you find there.

Data:

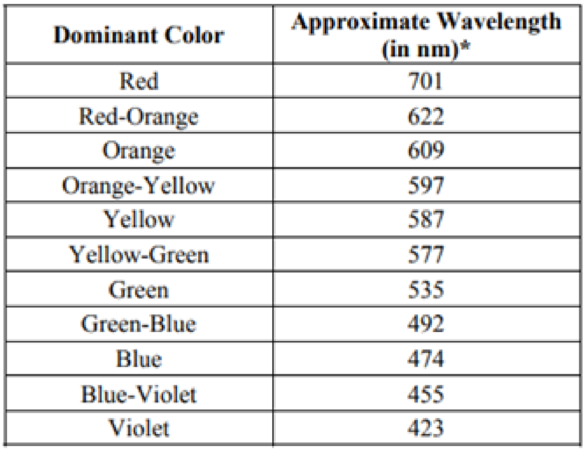
a) Complete the data tables below.

Cation in the salt	**Flame Color and Intensity**	**Unknowns**	**Flame Color(s) and Intensities**
Barium (Ba^{2+})		Unknown #1	
Calcium (Ca^{2+})		Unknown #2	
Copper (Cu^{2+})		Unknown #3	
Lithium (Li^{+})		Unknown #4	
Potassium (K^{+})		Unknown #5	
Sodium (Na^{+})		Unknown #6	
Strontium (Sr^{2+})			

b) Complete the chart below to identify each of the unknowns. You may identify them by symbol or by name.

Unknowns	**Identity or identities of unknown**
Unknown #1	
Unknown #2	
Unknown #3	
Unknown #4	
Unknown #5	
Unknown #6	

c) Select *one* of the seven known metal salts. Use the wavelength data for that color (given in the chart below) to calculate the frequency in Hz and the energy in eV of one photon of that color of light. Show all the necessary formulas, substitutions and work.



Dominant Color	Approximate Wavelength (in nm)*
Red	701
Red-Orange	622
Orange	609
Orange-Yellow	597
Yellow	587
Yellow-Green	577
Green	535
Green-Blue	492
Blue	474
Blue-Violet	455
Violet	423

d) Use appropriate atomic structure vocabulary and concepts to explain why an element produces a specific color of light during a flame test, in a well-written paragraph.

e) A student is testing a new unknown, Unknown #7, in a lit Bunsen burner flame. The flame is half green and half purple.

Is Unknown #7 an element, a compound, or a mixture? If Unknown #7 is a compound or a mixture, which elements (from this lab) are in it? Explain your answer using appropriate chemical vocabulary and concepts related elements, compounds, mixtures, and intensive physical properties.

2.06: Modern Atomic Theory and Quantum Numbers

1. (3.B, 3.F, 6.A) Compare Bohr’s atomic model with the quantum mechanical model.

2. (3.B, 6.A) Are the following sets of quantum numbers allowable? If they aren’t, state the error(s). (If a set isn’t allowable, that electron thus described doesn’t exist.)

a) $n = 1$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$

b) $n = 2$, $l = 3$, $m_l = -2$, $m_s = -\frac{1}{2}$

c) $n = 4$, $l = 1$, $m_l = -1$, $m_s = +1$

3. (3.B, 6.A) State the principal energy level and sublevel for each pair of quantum numbers.

	Principal Energy Level	Sublevel
a) $n = 4$, $l = 3$		
b) $n = 2$, $l = 0$		
c) $n = 3$, $l = 1$		
d) $n = 5$, $l = 2$		

4. (3.B, 6.A) Which of these electron transitions indicates a move between energy levels? Between sublevels? Is energy released or absorbed? An example is done for you.

a) From 1s to 2s: energy is absorbed, the electron is excited as it moves up energy levels

b) From 2p to 2s

c) From 3p to 2s

5. (3.A, 3.B) Use LASA as the backdrop of an analogy you can use to describe the structure of the electron cloud. State what part(s) of LASA are like the energy levels, sublevels, orbitals, and electrons. Then, evaluate: How good is this LASA analogy at describing electron cloud structure?

Lesson 2.06 Student Notes:

2.07: Electron Configurations and Orbital Notation

1. (6.D) Write the ground state electron configuration for an atom of each element:

a) Nitrogen

b) Titanium

c) Platinum

2. (3.B) Which element is denoted by each ground state electron configuration? What is its valence energy level and how many valence electrons are there?

Ground State Electron Configuration	Element symbol	Valence Energy Level	# Valence Electrons	Highest Energy Orbital
a) $1s^2\ 2s^2\ 2p^6\ 3s^2$				
b) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 4s^2\ 3d^{10}\ 4p^4$				
c) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^1$				

3. (3.B, 6.D) List an element whose highest energy orbital has the following *n* and *l* values. (There is more than one answer; just select one.) Then, write the ground state electron configuration for an atom of that element.

	Element Symbol	Ground State Electron Configuration
a) *n* = 2, *l* = 1		
b) *n* = 4, *l* = 2		
c) *n* = 5, *l* = 0		

4. (6.D) Write the ground state electron configuration for Ca^{2+}, Ne, and S^{2-}. Which species are isoelectronic to each other? Justify your answer.

5. (3.B) Identify each atom by adding up all of the electrons. (Assume all atoms are neutral.) Consult a periodic table for the atomic number. Then, tell if the atom is excited or if it is in the ground state.

a) $1s^2\ 3p^1$

b) $1s^2\ 2s^2\ 2p^3\ 3s^1$

6. (6.D) Write the noble gas core ground state electron configuration for:

a) Manganese

b) Palladium

c) Tantalum

d) Meitnerium

e) Radon (no, it isn't [Rn])

7. (3.B) Identify the element, then state the valence energy level and the number of valence electrons.

	Element	Valence Energy Level	# Valence Electrons
a) [Kr] $5s^2\ 4d^{10}\ 5p^6$			
b) [Ne] $3s^2\ 3p^1$			
c) [Xe] $6s^2\ 4f^1$			

8. (3.B, 6.D) Write the noble gas core electron configuration for each of the following elements:

a) Molybdenum

b) Silver

c) Gold

9. (3.B) Draw the unabbreviated orbital notation for each of the following atoms.

a) Silicon

b) Germanium

c) Tellurium

10. (3.B) Draw the noble gas core orbital notation for each atom.

a) Tungsten

b) Terbium

c) Plutonium

11. (3.B) For each set of quantum numbers, name the element that has *this specific* electron as the *highest* energy electron. Draw the ground state noble gas core electron configuration for this element and make this electron the last arrow you draw in the notation.)

	Element	Noble Gas Core Orbital Notation
a) $n = 4$, $l = 2$, $m_l = 0$, $m_s = -\frac{1}{2}$		
b) $n = 2$, $l = 1$, $m_l = 1$, $m = +\frac{1}{2}$		
c) $n = 3$, $l = 0$, $m_l = 0$, $m_s = -\frac{1}{2}$		

12. (3.B) Draw ground state noble gas core electron configurations for each of the following atoms and state the four quantum numbers for the last electron you draw in the notation.

a) Silicon

b) Arsenic

c) Iridium

Lesson 2.07 Student Notes:

2.08: Density and the Periodic Table Lab

Before we undertake our study of the periodic table, we would like for you to understand how the properties of the elements are used to help predict trends from actual data. We will use density as our (intensive) physical property of choice to analyze how you can use the periodic table in this way.

Problem*:*

Determine the densities of carbon, silicon, tin, and lead by measuring the mass and volume of a sample of each element. Then use this information to develop a claim for the numerical densities of germanium and flerovium in g/cm^3. This will help you understand how various elements on the periodic table are related in terms of their physical and chemical properties.

Materials:

- Carbon rods
- Silicon pieces
- Granular tin
- Lead shot
- Water
- 4 beakers (each 250 mL)
- Graduated cylinders (variety of sizes)
- Electronic balance
- Periodic table
- Weighing boats

Your group is designing and implementing a procedure to collect the appropriate data. To design your procedure and complete your objective, you will need to think about what type of data you need to collect, how you will collect the data, and how you will analyze the data.

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

To determine what type of data you need to collect, think about the following questions:

- What type of information will you need to collect during your investigation to determine the densities of carbon, silicon, tin, and lead?
- How will you use the densities of carbon, silicon, tin, and lead to predict the densities of the other two elements?

To determine how you will collect the data, think about the following questions:

- How will you reduce measurement error?
- How will you keep track of the data you collect?

To determine how you will analyze the data, think about the following questions:

- What calculations will you need to determine the densities of each element?
- How can you use mathematics to predict the densities of the other two elements based on the densities of carbon, silicon, tin and lead?
- What type of graph could you create to help make sense of your data?
- What variable will you graph against density – atomic mass, atomic number, or period number? Which scientific principles support your decision?

Once your data is collected, you will need to create and use a data table. All columns and rows should be labelled with a name and a unit, if applicable. (ex. Mass in g). Graph your data, title it descriptively, and label the axes and equation.

Written Report*:* Use the Google Docs template in BLEND.

EVERYONE (each individual person) must make substantial contributions to the report. If a person does NOT contribute to the report, as shown in the version history, that person will receive a zero on the lab.

2.09: The Periodic Table and Chemical Families

1. (5.B, 6.D) Write the ground state electron configuration (unabbreviated or noble gas core, your choice) of each element. State three physical and/or chemical properties of each element and the chemical family/families it belongs to. List how many valence electrons an atom of that element has and its valence energy level.

a) Magnesium

b) Chlorine

c) Titanium

e) Xenon

f) Francium

g) Uranium

2. (3.B, 5.B) Select three elements in the same group of the periodic table – the alkali metals, alkaline earth metals, halogens, or noble gases.

a) Write the ground state electron configuration for each element you select.

b) What similarities and differences do you observe between the configurations of the three elements?

c) Based on their electron configurations, explain why these three elements have similar physical and chemical properties.

3. (3.A, 3.B, 5.A) There are many historical and modern “alternative” versions of the periodic table. Why is Mendeleev’s periodic table given so much credit? Explain, making reference to physical and chemical properties of the elements somewhere in your answer.

4. (3.B, 3.F, 5.A) In the “modern” era, Glenn Seaborg is highlighted because he focused his life’s work on the lanthanides and actinides. Research Glenn Seaborg’s work. Briefly explain his work and explain why he placed the lanthanides and actinides together, in their own “block.”

Challenge Problem

5. (3.A, 3.B, 5.A) Practice Question #1 in this lesson alluded to the many “alternative” historical and modern periodic tables out there. Select one and research that periodic table.

a) Briefly explain the features of this “alternative” periodic table.

b) Analyze and evaluate this “alternative” periodic table. Is it competent at organizing the elements by their physical and chemical properties and how well does it do this?

6. (2.H, 2.I, 3.B, 3.F, 5.A) In the “modern” era, many elements have been discovered and synthesized by labs in Russia, Japan, and the United States, and their names reflect their countries of origin. Since these elements are difficult to isolate, we need to use the periodic table to infer their properties.

Select one of these elements and predict three of its physical and/or chemical properties based on its placement in the periodic table and its proximity to elements in the same block or group.

Lesson 2.09 Student Notes:

2.10: Periodic Trends

1. (3.B) Compare each element in each pair in terms of their Z_{eff} and their shielding effect. One is done for you.

Elements	Comparing Effective Nuclear Charge	Comparing Shielding Effect
Na and Mg	Mg has the greater Z_{eff}	Same shielding effect
P and As		
V and Cr		
Ge and Sn		

2. (3.B, 5.C) Of cesium, Cs, hafnium, Hf and gold, Au, which element has the smallest atomic radius? Explain your answer, using relevant concepts such as Coulombic attraction, effective nuclear charge, electron repulsion, and/or the shielding effect and their effect on the atomic radius.

3. (3.B, 5.C) Which element has the largest radius in the halogens group? Explain your answer, using relevant concepts such as Coulombic attraction, effective nuclear charge, electron repulsion, and/or the shielding effect and their effect on the atomic radius.

4. (3.B, 5.C) The two ions K^+ and Ca^{2+} each have 18 electrons surrounding the nucleus. Which would you expect to have the smaller radius? Explain your answer, using relevant concepts such as Coulombic attraction, effective nuclear charge, electron repulsion, and/or the shielding effect and their effect on the atomic radius.

5. (3.B, 5.C) Which ion would you expect to have the smaller radius, S^{2-}, Cl^- or Ca^{2+}? Explain your answer, using relevant concepts such as Coulombic attraction, effective nuclear charge, electron repulsion, and/or the shielding effect and their effect on the atomic radius.

6. (3.B, 5.C) Use scientific principles (such as are explained here) to explain each of the following observations.

a) The first ionization energy of Fr is the lowest in the alkali metals group.

b) The first ionization energy of Ne is the largest in period 2.

c) The first ionization energy of Ga is lower than the first ionization energy of Ca.

7. (2.H, 2.I, 3.B, 5.C) The following data list the first, second, third and fourth ionization energies for a mystery metal, M. (kJ/mol = unit for ionization energy). What could be a possible charge for an ion of metal M? Explain your answer using principles of atomic structure.

Element	First IE (kJ/mol)	Second IE (kJ/mol)	Third IE (kJ/mol)	Fourth IE (kJ/mol)
M	577	1820	2740	11600

8. (3.B, 5.C) Explain why the electronegativity of Pb is the smallest in its group, using principles of atomic structure.

9. (3.B, 5.C) Explain why the electronegativity of I is the largest in its period, using principles of atomic structure.

10. (3.B, 5.C) Briefly explain why the electron affinity of F is the largest and most negative of all the elements in its period.

Challenge Problems:

11. (3.B, 5.C) Which of the following cations is least likely to form and why, Sr^{2+} or K^{2+}? Use principles of atomic structure to explain this statement.

Lesson 2.10 Student Notes:

Unit 3: Chemical Bonding

3.01: Ionic and Metallic Bonding Concepts

1. (3.A, 3.B, 7.D) Evaluate the “sea of electrons” model: is it strong enough to explain the malleability, ductility, luster, and conductivity of metals? (Hint: the central idea is that the mobile valence electrons float freely throughout the flexible structure, but you must elaborate on this idea in your explanations.)

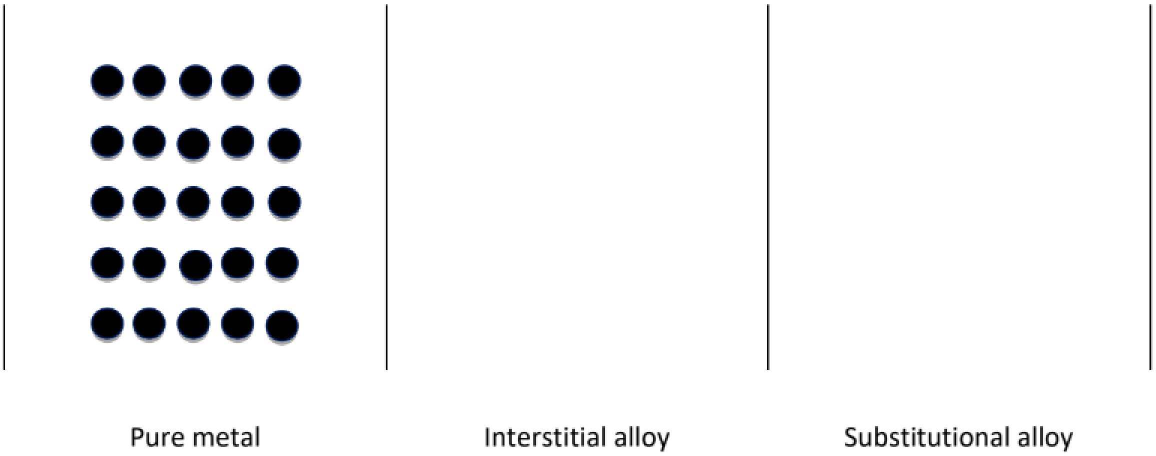
a) Malleability:

b) Ductility:

c) Luster:

d) Conductivity:

2. (7.D) Draw a particle diagram illustrating your understanding of interstitial and substitutional alloy structure. Use another (different) symbol for atoms of the second metal. One is done for you. (For clarity, you do not have to draw the mobile electrons.)



3. (7.C) Draw the Lewis valence electron-dot structure for each element.

Strontium Boron Carbon Oxygen Iodine

4. (7.C) Draw the Lewis electron-dot structure for each formula unit, showing work.

$AlBr_3$

K_2O

between aluminum and phosphorus

between calcium and nitrogen

5. (2.H, 2.I, 3.A, 7.D) Use principles of atomic structure and/or chemical bonding to explain the following observations. Discuss both substances in your response.

a) The melting point of KCl is lower than the melting point of CaS, even though both are ionic compounds.

b) Mg, a metal, is malleable and ductile, but MgO, an ionic compound, is brittle.

c) Molten NaCl, an ionic compound, conducts electricity better than solid NaCl.

d) $MgCl_{2\ (s)}$ has a higher melting point than $SrCl_{2\ (s)}$.

Lesson 3.01 Student Notes:

Common Polyatomic Ions Reference

The following ions, both formula AND charge WITH name, must be memorized **if bolded**. Monatomic ions, such as calcium, fluoride, etc. are *not* here but still need to be memorized.

Cations

+1 charge

Ammonium	**NH_4^+**
Silver	**Ag^+**

+2 charge

Zinc	**Zn^{2+}**

Variable Charge

Iron (II) ("ferrous")	**Fe^{2+}**
Iron (III) ("ferric")	**Fe^{3+}**
Copper (I) ("cuprous")	**Cu^+**
Copper (II) ("cupric")	**Cu^{2+}**
Lead (II) ("plumbous")	**Pb^{2+}**
Lead (IV) ("plumbic")	**Pb^{4+}**
Tin (II) ("stannous")	**Sn^{2+}**
Tin (IV) ("stannic")	**Sn^{4+}**

Anions

-1 charge

Hydride	**H^-**
Nitrite	**NO_2^-**
Nitrate	**NO_3^-**
Hypochlorite	**ClO^-**
Chlorite	**ClO_2^-**
Chlorate	**ClO_3^-**
Perchlorate	**ClO_4^-**
Bromate	**BrO_3^-**
Iodate	**IO_3^-**
Hydroxide	**OH^-**
Cyanide	CN^-
Acetate	**$C_2H_3O_2^-$ OR CH_3COO^-**
Permanganate	MnO_4^-
Hydrogen carbonate	**HCO_3^-**
Hydrogen sulfate	**HSO_4^-**
Hydrogen sulfite	**HSO_3^-**

-2 charge

Carbonate	**CO_3^{2-}**
Hydrogen phosphate	**HPO_4^{2-}**
Sulfite	**SO_3^{2-}**
Sulfate	**SO_4^{2-}**
Chromate	CrO_4^{2-}
Dichromate	$Cr_2O_7^{2-}$
Thiosulfate	$S_2O_3^{2-}$
Oxalate	**$C_2O_4^{2-}$**
Peroxide	**O_2^{2-}**

-3 charge

Phosphite	**PO_3^{3-}**
Phosphate	**PO_4^{3-}**

3.02: Ionic Compound Nomenclature (No Transition Metals)

1. (5.B) What are the expected charges for all elements in each group of the periodic table?

Group 1 =

Group 2 =

Group 13 =

Group 15 =

Group 16 =

Group 17 =

2. (7.A, 7.B) If given the name, write the corresponding IUPAC formula, and vice versa.

SrSe

Rubidium nitride

K_2Te

Na_2S

Aluminum oxide

Al_2S_3

3. (7.A, 7.B) For each given name, write its IUPAC chemical formula, and vice versa.

Ammonium chloride

Cesium chlorite

$(NH_4)_3PO_4$

Magnesium dichromate

$Al_2(HPO_4)_3$

Calcium oxalate

Na_2O_2

$Al(HSO_4)_3$

Lithium thiosulfate

Lesson 3.02 Student Notes:

3.03: Ionic Compound Nomenclature (Including Transition Metals)

1. (7.A, 7.B) For each given name, write its IUPAC chemical formula, and vice versa.

Cr_2O_3

Silver iodide

NiN

Cobalt (III) sulfide

Cu_3P_2

Lead (II) oxide

Stannic fluoride

Cuprous nitride

PbF_2 (use old name)

2. (7.A, 7.B) For each given name, write its IUPAC chemical formula, and vice versa.

Iron (III) peroxide

$AuMnO_4$

Ammonium chlorate

$Pb(ClO_2)_2$

Copper (II) hydroxide

Cupric chromate

$Sn(HSO_4)_2$

$Ag_2C_2O_4$

3. (7.A, 7.B) This question draws from 3.02 and 3.03. Write the IUPAC name for each formula, or vice versa.

strontium fluoride ____________________

zinc nitride ____________________

lead (II) oxalate ____________________

ammonium oxide ____________________

copper(I) thiosulfate ____________________

cuprous peroxide ____________________

K_2S ____________________

$NaNO_3$ ____________________

Ag_2CO_3 ____________________

$NH_4C_2H_3O_2$ ____________________

$Ni(CN)_2$ ____________________

$Fe(NO_2)_2$ ____________________

ferrous sulfate ____________________

iron (III) chromate ____________________

plumbic hydroxide ____________________

gallium (III) acetate ____________________

zinc dichromate ____________________

SnF_2 ____________________

Al_2S_3 ____________________

$(NH_4)_2SO_3$ ____________________

$Ba(HSO_3)_2$ ____________________

$CuSO_3$ ____________________

Challenge Question:

4. (7.A, 7.C) Determine the charge of X in each ionic compound listed below. Give one possible identity for X, draw its Lewis valence electron-dot diagram, and write out the noble gas core electron configuration for the ion of the element you selected.

X_2O

X_3N_2

CaX_2

XBr_3

Lesson 3.03 Student Notes:

3.04: Covalent Bonding Concepts and Covalent / Acids Nomenclature

1. (7.A, 7.B) Write the IUPAC formula or the name for each covalent compound.

carbon dioxide		BrF_3	
diphosphorus pentasulfide		CS_2	
dinitrogen trioxide		CBr_4	
sulfur hexachloride		XeS_3	

2. (7.A, 7.B) Write the IUPAC name or formula for each acid.

hydrobromic acid		H_3PO_3	
perchloric acid		HI	
hypochlorous acid		$HBrO_3$	
oxalic acid		H_2CO_3	
sulfuric acid		$HClO_3$	

3. (7.A, 7.B) Write the IUPAC name or formula for each compound. Warning - there are mixed types!

iron (II) oxalate		NaH	
iodic acid		ClO_2	
calcium nitrate		H_2SO_4	
sodium sulfide		$FeCO_3$	
arsenic trifluoride		$NH_4C_2H_3O_2$	

calcium sulfate ____________

stannic fluoride ____________

tin (II) nitrate ____________

potassium chlorate ____________

hydroxic acid ____________

bromine pentaiodide ____________

Cuprous hydrogen phosphate ____________

stannous acetate ____________

IF_7 ____________

$H_2S_2O_3$ ____________

$HClO_2$ ____________

SF_4 ____________

$NiSO_4$ ____________

$Cr(ClO_2)_3$ ____________

ClO ____________

XeF_4 ____________

Challenge Question**:**

4. (7.A, 7.B) In the space below make a naming compound flowchart including ionic, covalent, and acid nomenclature. If you need more space, use the Lesson 3.04 student note page.

Lesson 3.04 Student Notes:

3.05: Lewis Structures (including Exceptions)

1. (7.C) Draw the Lewis valence electron dot structure for each element.

Li Ca C Br He Kr

2. (7.C) Draw the Lewis structure for each molecule.

H_2S PBr_3 SO_3

SO_2 HCN CO

C_3H_8 C_2H_4 NI_3

3. (7.C) Draw the Lewis structure for each molecule.

$BeCl_2$ BF_3

4. (7.C) Draw the Lewis structure for each molecule.

$XeCl_2$

SF_4

SF_6

BrF_5

KrF_2

PBr_5

5. (7.C) Draw the Lewis structure for each ion.

OH^-

CO_3^{2-}

NH_4^+

ClO_2 -

ClO_3^-

NO_3^-

CN^-

SO_3^{2-}

SO_4^{2-}

6. (7.C) Draw all the resonance structures for carbonate.

7. (7.C) Mixed practice: Draw the Lewis structure for each molecule or ion.

PCl_5

XeF_2

PO_3^{3-}

NF_3

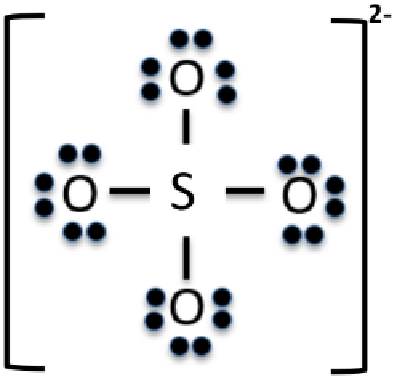
BI_3

NO_2^-

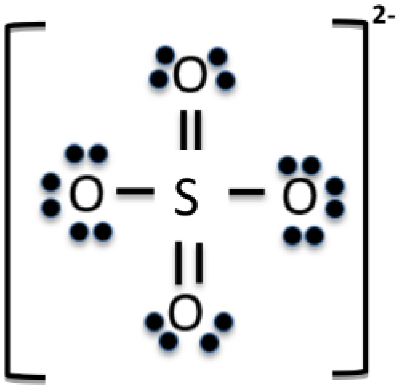
Lesson 3.05 Student Notes:

3.06: Formal Charge and VSEPR, including Organic Molecules

1. (7.C) Calculate all formal charges for the sulfate structures and determine which is the best structure.



a)



b)

2. (7.C) Draw three different and valid Lewis structures for xenon trioxide. Mark the formal charge by each atom and determine which is the best structure based on formal charges.

If your teacher so chooses, they may have you complete the following lab.

MODEL 1: ***Molecule Shapes*** **Simulation**
(http://phet.colorado.edu/en/simulation/molecule-shapes)

PART I: ELECTRON DOMAINS

1. Explore the *Model* screen of the simulation. As you explore, answer the following questions.

a) How does adding an **atom** (the spheres under bonding to right) affect the position of existing atoms or lone pairs?

b) How does adding a **lone pair** affect the position of existing atoms and lone pairs?

2. Is the effect of adding bonded **atoms** and **lone pairs** to the central atom similar? Explain why this could be the case.

> We can think of a bond or a lone pair of electrons as a "domain" of electrons. Single bonds, double bonds, and triple bonds each count as one domain.

3. How do the electrons in bonds (bonding domains) differ from lone pairs (non-bonding domains)?

4. What happens to the **bond angle** when you add or remove an electron domain?

5. Can you force the atoms into new configurations by pushing atoms around? What does this suggest about the configuration of atoms in real molecules?

6. What is the difference between *Electron Geometry* and *Molecule Geometry*?

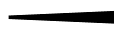
7. In one or two grammatically correct sentences, write a definition for the term *Molecule Geometry*.

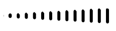
PART 2: DRAWING MOLECULES TO SHOW 3-DIMENSIONALITY

MODEL 2:

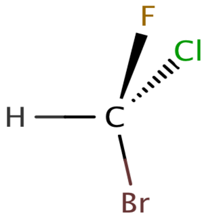
Line, Wedge and Dash Drawings

Line: In the plane of the paper: ——

Wedge: Coming forward, in front of the plane of the paper: 

Dash: Going backward, behind the plane of the paper: 

8. Where is each of the 5 atoms in the molecule CHFClBr?
   In the plane of the paper ______ ______ ______
   In front of the plane of the paper ______
   Behind the plane of the paper ______

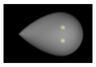


9. Using the *Model* screen, add bonding domains (●) to the central atom (○). Using lines, wedges, and dashes from Model 2, draw each molecule's shape.

Bonding Domains Around Central Atom	Drawing of Shape	Electron Geometry	Bond Angles

2	● — ○ — ●	Linear	180º
3	○		
4	○		
5	○		
6	○		

10. In the *Model* screen, build a molecule with 5 atoms attached to the central atom. Look at the molecule geometry and electron geometry. **Predict** what will happen to the molecule geometry as you replace atoms with lone pairs.

11. In the following table draw the **molecule geometry** using the bonding domains (●) and lone pairs ()to the central atom (○). Using lines, wedges, and dashes, draw each molecule's shape.

Predict First, Then Compare with the Simulation				
Number of Domains Around Central Atom	1 Lone Pair	2 Lone Pairs	3 Lone Pairs	4 Lone Pairs
3				

4			▨	▨
5				▨
6				

PART 3: COMPARING MODEL VS. REAL MOLECULES

12. Explore the *Real Molecules* screen.

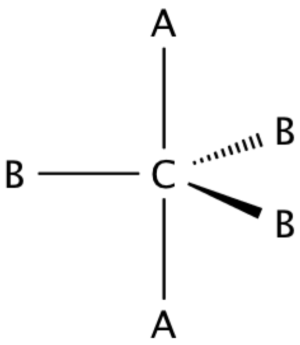
a) List the molecules that show a **difference in bond angle** between “Real” and “Model”. Note: differences in bond angle may be small.

Molecule	Number of Lone Pair Domains

b) What do all the molecules in the table have in common?

c) What trend do you observe that distinguishes lone pairs from bonding domains?

13. Use the simulation to build a system with 5 domains. This is called a **trigonal bipyramidal structure**. The two different sites in a trigonal bipyramid are labeled as A and B in the drawing to the right.



a) Each A atom is adjacent (next to) to 3 B atoms. What is the A-C-B bond angle?

b) Each B atom is adjacent to 2 A atoms and 2 B atoms. What is the B-C-B bond angle?

c) In a system with 4 atoms and 1 lone pair, predict whether the lone pair will be in a B site or an A site? Explain.

d) Examine the molecule SF_4 in the Real Molecules screen to check your prediction from question c). Which interactions (forces) are more important in determining where the lone pair will go?

Challenge Question:

Imagine a molecule with 7 electron domains.

The geometry is called **pentagonal bipyramidal**. This has a lot of similarities to a 6-coordinate molecule except there are 5 domains in one plane. Predict the following based on this information:

a) Draw the structure of a molecule with 7 bonding domains.

b) List all the bond angles possible between adjacent atoms.

c) Predict the electron and molecule geometry for a molecule with 6 bonding domains and a single lone pair.

d) Predict the electron and molecule geometry for a molecule with 5 bonding domains and two lone pairs.

Herzog, Tiimothy, and Moore, Emily. "Molecule-Shapes-Student-Handout." PhET, 2 Nov. 2014. Accessed at: https://phet.colorado.edu/en/contributions/view/3947. Licensed under CC BY 4.0: https://creativecommons.org/licenses/by/4.0/legalcode.

VSEPR (Valence Shell Electron Pair Repulsion) Theory

3. (7.C, 7.E) Draw the Lewis structure for each molecule, using formal charges to determine the most stable structure. Determine its electron-dot geometry, molecular shape and bond angle(s).

NF_3	SO_2	ICl_3	SbF_5
H_2Se	BF_3	PCl_3	CCl_4

O_3
(put O's in a *line*!)

SCl_4

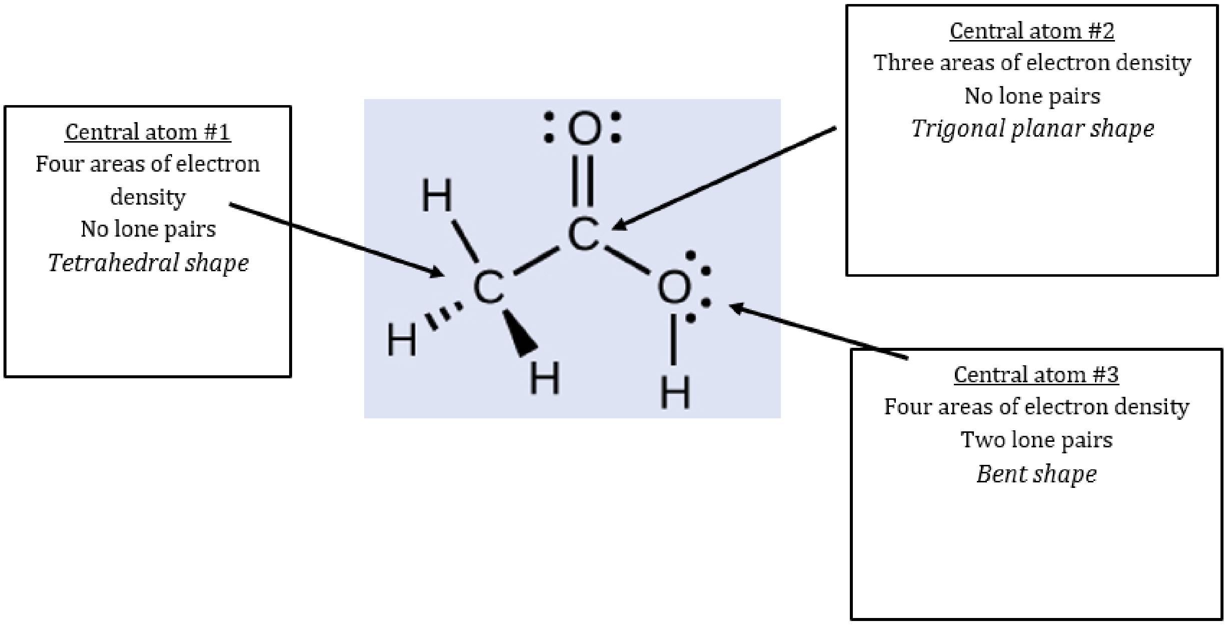
$KrCl_2$

$XeBr_4$

Challenge Question

4. (7.E) Read the text below and answer the question for the methionine molecule.

Organic molecules – those containing primarily C and H – have more complex Lewis structures. Different parts of the same molecule can have different VSEPR shapes since there are multiple central atoms. You need only define the VSEPR molecular shape around each central atom. Examine the structures below for examples:[4]



Central atom #1
Four areas of electron density
No lone pairs
Tetrahedral shape

Central atom #2
Three areas of electron density
No lone pairs
Trigonal planar shape

Central atom #3
Four areas of electron density
Two lone pairs
Bent shape

(You can also say the O double bonded to C has a trigonal planar geometry, since there are three areas of electron density around it.)

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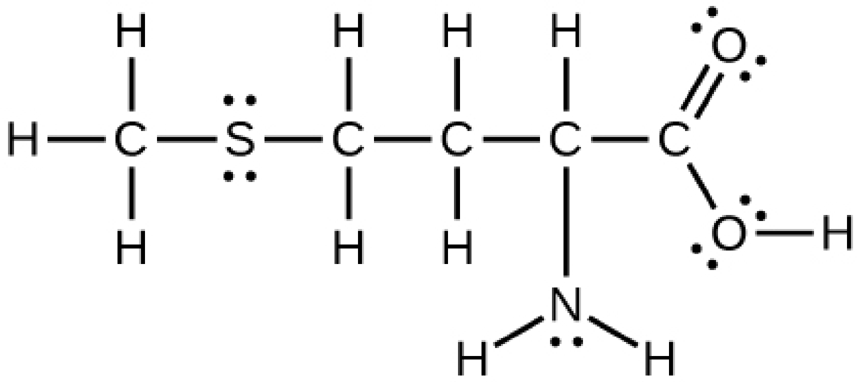
[4] "Acetic Acid." *8.2 Hybrid Atomic Orbitals,* , Rice University, https://opentextbc.ca/chemistry/chapter/8-2-hybrid-atomic-orbitals/. Licensed under CC BY 4.0: https://creativecommons.org/licenses/by/4.0/legalcode

a) Determine all the VSEPR shapes possible around each atom in the methionine molecule below (all atoms, except for the hydrogens).[5]



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[5] "Methionine." *8.2 Hybrid Atomic Orbitals,* Rice University, https://opentextbc.ca/chemistry/chapter/8-2-hybrid-atomic-orbitals/. Licensed under CC BY 4.0: https://creativecommons.org/licenses/by/4.0/legalcode

3.06 Student Notes:

3.07: Bond and Molecular Polarity

1. (7.C, 7.E) Draw the Lewis structure of each molecule or ion. Take formal charge into account and draw all resonance structures if applicable. Don't forget to draw all the lone pair electrons around the atoms that have them. Then, identify the structure's VSEPR molecular shape.

hydrocyanic acid, HCN

$POCl_3$ (Cl = chlorine)

nitrite ion

PCl_2^{-1} (Cl = chlorine)

iodine trichloride

chlorine pentafluoride

xenon tetrabromide

sulfur trioxide

bromine pentafluoride

krypton dichloride

2. (3.B) Use your knowledge of electronegativity trends to assign **σ**+ and **σ**- to the dipoles on all the polar bonds. If the bond is nonpolar, leave that question blank, or state “nonpolar”.

a) O = O

b) F – Cl

c) C - N

3. Go back to question #1 and identify each molecule or ion as polar or nonpolar based on symmetry or asymmetry in their VSEPR molecular shape.

Lesson 3.07 Student Notes:

Unit 4: Chemical Reactions

4.01: Chemical Reactions Concepts

1. (8.E) How many of each type of atom are in each formula? Note: Charges don't change how you count atoms.

$3\ CaCO_3$ $Ca(NO_3)_2$ $2\ Cr(C_2O_4)_3^{3-}$

2. (8.E) Draw a particle diagram depicting the change in the balanced chemical equation. Use whatever symbols and color scheme you like.

$$C_2H_{4\ (g)} + 3\ O_{2\ (g)} \rightarrow 2\ CO_{2\ (g)} + 2\ H_2O_{\ (g)}$$

Provide work – not just the correct coefficients - for the questions that follow.

3. (8.E) Balance the following chemical equation. States have been omitted for clarity.

$$\_\_\ N_2 + \_\_\ F_2 \rightarrow \_\_\ NF_3$$

4. (8.E) Balance the following chemical equation. States have been omitted for clarity.

$\_\_\ Na_3PO_4 \quad + \quad \_\_\ KOH \quad \rightarrow \quad \_\_\ NaOH \quad + \quad \_\_\ K_3PO_4$

5. (8.E) Balance the following chemical equation. States have been omitted for clarity.

$\_\_\ Fe(OH)_3 \quad \rightarrow \quad \_\_\ Fe_2O_3 \quad + \quad \_\_\ H_2O$

6. (8.E, 8.F) Balance each chemical equation and identify the type of reaction.

$\_\_\ C_3H_6 \quad + \quad \_\_\ O_2 \quad \rightarrow \quad \_\_\ CO_2 \quad + \quad \_\_\ H_2O$

7. (8.E, 8.F) Balance each chemical equation and identify the type of reaction.

$\_\_\ NH_3 \quad + \quad \_\_\ O_2 \quad \rightarrow \quad \_\_\ NO \quad + \quad \_\_\ H_2O$

8. (8.E, 8.F) Balance each chemical equation and identify the type of reaction.

__ NaI + __ $Pb(SO_4)_2$ → __ PbI_4 + __ Na_2SO_4

9. (8.E, 8.F) Balance each chemical equation and identify the type of reaction.

__ C_5H_{10} + __ O_2 → __ CO_2 + __ H_2O

10. (8.E, 8.F) Balance each chemical equation and identify the type of reaction.

__ Ba + __ H_2O → __ $Ba(OH)_2$ + __ H_2

Challenge Question:

11. (3.B, 8.E) Balance the following chemical equation. If you want to challenge yourself further, try to solve it using linear matrices.

__ A_2C_3 + __ B_2CD_3 → __ A_2 + __ A_2B_4 + __ C_9D_{10}

Lesson 4.01 Student Notes:

4.02: Redox Concepts

1. (2.G, 3.B) Assign the oxidation number to the atom in **bold** print.

a) **N** in NO_3^-

c) **Cr** in $Cr_2O_7^{2-}$

b) **Cl** in $NaClO_4$

d) **Pb** in $Pb(OH)^+$

2. (2.G, 3.B) Assign oxidation numbers to *every* element in each molecule or ion.

a) sodium hydride

c) H_2O_2

e) chlorate ion

b) oxalate ion

d) carbonic acid

f) aluminum phosphate

3. (2.G, 3.B) Identify each of the following changes as either an oxidation half-reaction or a reduction half-reaction. The states are omitted for clarity.

a) $I_2 + 2\,e^- \rightarrow 2\,I^-$

c) $Fe^{2+} \rightarrow Fe^{3+} + 1\,e^-$

b) $K \rightarrow K^+ + 1\,e^-$

d) $Ag^+ + 1\,e^- \rightarrow Ag$

4. (2.G, 3.B) Identify what is oxidized, what is reduced, the oxidizing agent, and the reducing agent for each reaction. States have been omitted for clarity.

a) $Fe + CuCl_2 \rightarrow Cu + FeCl_2$

b) $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

c) $Cu + 4\ HNO_3 \rightarrow Cu(NO_3)_2 + 2\ NO_2 + 2\ H_2O$

d) $2\ NaClO_3 \rightarrow 2\ NaCl + 3\ O_2$

e) $2\ Mg + O_2 \rightarrow 2\ MgO$

5. (2.G, 3.B, 8.F) The following reactions may or may not be redox. Identify all the oxidation numbers for each type of atom, then determine if the overall reaction is redox or not redox. **If it is redox, the oxidation numbers for the same type of atom will change as you read the reaction from left to right.** States have been omitted for clarity.

a) $2\ KBr + F_2 \rightarrow 2\ KF + Br_2$

b) $H_2 + O_2 \rightarrow H_2O_2$

c) $(NH_4)_2CO_3 \rightarrow 2\ NH_3 + CO_2 + H_2O$

d) $Pb(NO_3)_2 + NaI \rightarrow PbI_2 + NaNO_3$

e) $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

Challenge Problems:

6. (3.B, 8.E) Balancing a redox reaction requires that you balance both the atoms as well as the charges. Research “balancing redox reactions using the half-reaction method,” and use this method to balance the following redox reactions.

a) $Mn^{2+} + BiO_3^- \rightarrow MnO_4^- + Bi^{2+}$ (in acidic solution)

b) $H_2S + NO_3^- \rightarrow S + NO$ (in basic solution)

Lesson 4.02 Student Notes:

4.03: Reaction Prediction Concepts

1. (3.B, 7.B) List the charge of elements in each group of the periodic table. Identify the chemical family of that group, if relevant.

a) Group 1A

b) Group 2A

c) Group 13A

d) Group 15A

e) Group 16A

f) Group 17A

2. (3.B, 7.B) Write the chemical formula of each compound.

a) sulfuric acid

b) calcium nitrate

c) dinitrogen trioxide

d) aluminum nitride

e) copper (II) hydroxide

f) plumbous sulfide

(2.H, 2.I, 3.B, 7.B, 8.E, 8.F) Predict and write balanced chemical equations for the following reactions.

3. Sulfur dioxide gas passes over a strip of copper (II) oxide.

4. Magnesium oxide powder is decomposed over very high heat.

5. A chunk of silicon is placed into a tank of oxygen gas, then sparked to cause a reaction.

6. Molten sodium chloride is decomposed.

7. Sodium oxide powder reacts with carbon dioxide gas.

8. Potassium chloride powder burns in oxygen gas.

9. Sodium chlorate powder is heated strongly.

10. Gaseous ethane, C_2H_6, burns in oxygen gas.

11. Liquid methanol, CH_3OH, burns in oxygen gas.

12. Calcium metal burns in oxygen gas.

13. Aluminum foil burns in oxygen gas.

14. (2.H, 2.I, 3.B, 7.B, 8.E, 8.F) Predict and write the balanced molecular equation and net ionic equation for the reaction between the following pairs of reactants – if they occur. If they don't, just write N.R.

a) Fluorine gas bubbles into a rubidium chloride solution.

b) Bromine liquid is gently mixed with aqueous lithium iodide.

c) Chlorine gas is passed into a potassium bromide solution.

15. (3.B) For one of the reactions in the previous question – you may choose - draw a labelled particle diagram of the beaker after the change is complete. Solvate each spectator ion with water molecules with the correct orientation. You may assume there is no excess reactant present.

16. (2.H, 2.I, 3.B, 7.B, 8.E, 8.F) Predict and write the balanced molecular equation and net ionic equation for the reaction between the following pairs of reactants – if they occur. If they don't, just write N.R.

a) A lump of chromium is dropped into a lead (II) chlorate solution. (Cr will have +3 charge)

b) Aqueous gold (I) chloride and a piece of magnesium ribbon are mixed together.

c) Barium metal is mixed with aqueous copper (II) bromide.

17. (3.B) For *one* of the reactions in the previous question – you may choose - draw a labelled particle diagram of the beaker after the change is complete. You must include water molecules and the spectator ion. You may assume there is no excess reactant present.

18. (2.H, 2.I, 3.B, 7.B, 8.E, 8.F) Predict and write the balanced net ionic chemical equation for each change.

a) Potassium metal is dropped in water.

b) Barium metal is dropped in water.

19. (3.B) For *one* of the reactions in the previous question – you may choose - draw a labelled particle diagram of the beaker after the change is complete. You must include water molecules and the spectator ion. You may assume there is no excess reactant present.

Lesson 4.03 Student Notes

4.04: Solubility Rules and Double Replacement Reactions

SOLUBILITY RULES: Rule #1 takes priority over rules #2-6.

1. All alkali metal (Group IA) and NH_4^+ (ammonium) salts are ***soluble.***
2. All Cl^-, Br^-, I^- salts are ***soluble***, *except for Ag^+, Hg_2^{+2}, Pb^{+2} salts. (Hg_2^{+2} is mercury (I) ion.)
3. All F^- salts are ***soluble,*** *except for Group IIA salts.
4. All NO_3^-, ClO_3^-, ClO_4^-, and CH_3COO^- / $C_2H_3O_2^-$ salts are ***soluble.***
5. All SO_4^{-2} salts are ***soluble,*** *except for Ca^{2+}, Sr^{+2}, Ba^{+2}, Ag^+, Hg_2^{+2}, Pb^{+2} salts.
6. All CO_3^{-2}, PO_4^{-3}, $C_2O_4^{-2}$, CrO_4^{-2}, OH^- salts and and all other salts are ***INSOLUBLE*** *except for hydroxide salts of Ca^{2+}, Sr^{+2}, Ba^{+2}.

1. (7.B, 10.B) Write the IUPAC chemical formula for each species. State if it is soluble or insoluble in water.

a) sodium iodide
b) ammonium sulfide
c) calcium fluoride
d) lead (II) bromide
e) aluminum acetate
f) aluminum fluoride
g) mercury (I) sulfate
h) copper (II) sulfate
i) lithium carbonate
j) tin (II) hydroxide

2. (3.B, 10.A) For *any* one of the soluble species and one of the insoluble species in #1, draw a labelled particle diagram showing the particles in water. Show the hydration of all aqueous ions and the correct ratio of ions present in the beaker.

Soluble species:

Insoluble species:

3. (3.B, 10.B) Which of the following ionic compounds in each set has the highest solubility in water? (Hint: this means the soluble one!)

a) calcium carbonate, lithium nitrate, mercury (I) chloride

b) iron (III) sulfate, copper (II) sulfide, magnesium oxide

c) manganese (II) chromate, silver iodide, silver chlorate

4. (3.B, 7.B, 8.E, 10.B) Identify the following compounds as either electrolytes or nonelectrolytes based on their solubility in water. If the solution is electrolytic, write a dissociation equation to show how the ions form. If the compound is a nonelectrolyte, leave it blank.

a) ammonium chromate

b) barium sulfate

c) sodium phosphate

5. (3.B, 7.B, 8.E, 8.F, 10.B) Predict and write the balanced molecular equation for *each* mixture, a) through e), including all state signifiers such as (s) or (aq).

If the mixture forms a precipitate, go further, and write the total ionic equation, the net ionic equation, and the two spectator ions. All state signifiers must be present and all ions must have correct charges.

a) Solutions of potassium phosphate and iron (II) nitrate are mixed.

b) Solutions of beryllium acetate and copper (I) sulfate are mixed.

c) Nickel (III) nitrate and rubidium sulfate solutions are poured into the same beaker.

d) Cobalt (III) chlorate solution is poured into aqueous potassium chromate.

e) Barium hydroxide solution is added to a manganese (II) nitrate solution.

6. (3.B, 10.A) Draw a particle representation of the beaker, its ions, and its precipitate (if any) for one pairing from a) – e) that produced precipitate and one pairing that produced all spectators. All aqueous ions must be hydrated. All ions must be shown in the correct ratio.

Precipitate produced:	All spectators produced:

7. (8.E) Draw a particle diagram showing the products of an acid-base neutralization reaction, with excess acid left over. You do not have to show the hydration of any ions, but you do have to include all spectator ions. You can use the chemical equation:
   $HCl_{(aq)}$ + $NaOH_{(aq)}$ → $NaCl_{(aq)}$ + $H_2O_{(l)}$ as a basis.

8. (7.B, 8.E) Write the balanced, molecular, neutralization equation for each reaction. Include all states. Assume that ionization is complete for all neutralizations. Include all states!

a) Hydrobromic acid and sodium hydroxide

b) Hydroiodic acid and calcium hydroxide

c) Sulfuric acid and rubidium hydroxide

Lesson 4.04 Student Notes:

4.05: Chemical Reactions Lab

Problem: Observe some chemical reactions and identify reactants and products of those reactions. Classify the reactions as synthesis, decomposition, single replacement, or double replacement, and write balanced chemical equations.

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

Materials:

Aluminum foil
Magnesium ribbon
Solid copper (II) carbonate
Hydrochloric acid solution
Copper (II) sulfate solution
*Isopropyl alcohol
*Large water jug
*Calcium metal
*Petri dish
*Alligator clips
*9V battery

Lead (II) nitrate solution
Potassium iodide solution
Sodium bicarbonate solution
Spot (well) plate
Bunsen burner
Crucible tongs
Spatula

Test tubes w/ rack
Test tube holder
Matches
Utility clamp
10 mL graduated cylinder
pH paper

*for optional stations or demos

Procedure

When you first get to a station (numbered below, #1-6), note the color and appearance of all reactants. Perform the chemical change in each step. As the chemical change occurs, note all color, appearance, and energy changes. Some stations have additional information. Please read all information carefully and follow all instructions.

Place all wastes in the beaker provided at each station, not in the sink. Used paper towels go into the trash. *Clean up the station before you leave!*

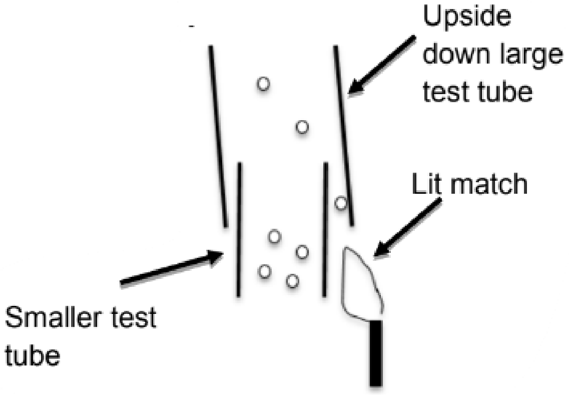
1. Aluminum foil (1 small piece) is added to 1 inch of copper (II) sulfate solution in a test tube.

2. Magnesium ribbon (1 inch) is heated in a Bunsen burner flame for 1-2 minutes. Use the crucible tongs to hold the ribbon.

READ BEFORE TESTING: Your teacher may choose to supervise this station. A very bright light will appear. *Do not look in to or at the light!*

3. Into a small diameter test tube, a 1-inch piece of magnesium ribbon is dropped into approximately one inch of hydrochloric acid. Quickly invert and cover the small test tube with the large test tube. Test for hydrogen, as described below.

READ BEFORE TESTING: Testing for hydrogen will require two students. One student lights a match. The other student will hold the smaller test tube in one hand and place the opening of the large test tube over the smaller test tube opening. Raise the large test tube so the opening of both test tubes line up. The large test tube should be tilted slightly open. Put the lit match close to the gap between the two test tubes. If hydrogen is present, you will *hear* something!



4. Copper (II) carbonate (1 tip full of a scoopula) is heated strongly in a large test tube over a Bunsen burner flame for about 2 minutes. Test for carbon dioxide, as described below.

READ BEFORE TESTING: Test for carbon dioxide by inserting a lit match into the test tube as far as you can go. Drop the match but watch it carefully as it falls down. If carbon dioxide is present, it will extinguish the flame on the way down (and not when it hits the bottom).

5. Potassium iodide (5 drops) and lead (II) nitrate (5 drops) solutions are mixed in a well plate.

6. Sodium bicarbonate solution (about 1 inch, in a large test tube) is mixed with hydrochloric acid.

READ BEFORE TESTING: Test each reactant with pH paper BEFORE you start the chemical change. Test the mixture with fresh pH paper AFTER the chemical change is complete.

Test for carbon dioxide by inserting a lit match into the test tube as far as you can go. Drop the match but watch it carefully as it falls. If carbon dioxide is present, it will extinguish the flame on the way down (and not when it hits the liquid).

Optional Demonstrations or Stations (at your teacher's discretion):

7. "Whoosh Bottle" Demonstration: Isopropyl alcohol, $C_3H_8O_{(aq)}$, is ignited with a flame.

8. A small sample of calcium metal is dropped into water (about 1 inch in a test tube).

READ BEFORE TESTING: Test the water with pH paper BEFORE you start the chemical change. Test for hydrogen as in station #3. Then, test the mixture with fresh pH paper AFTER the chemical change is complete. Any pH paper change is due to the formation of hydroxide ions.

9. Electrolysis of Water: Put a little deionized water in a Petri dish. Attach one mechanical pencil lead to each alligator clip. Hook the other end of both clips to the ports of a 9V battery. Add universal indicator to the water in the Petri dish. Insert both pencil leads into the water. The battery provides energy like a Bunsen burner would. (The pencil leads are not reactants.)

Observations Table (Only complete #7-9 if your teacher chose to have you do these stations.)

Reaction	Observations of Reactants Before Reaction	Observations During/After Reaction
1		
2		

3		
4		
Reaction	Observations of Reactants Before Reaction	Observations During/After Reaction
5		
6		
7		
8		
9		

Analysis: (Only complete #7-9 if your teacher chose to have you do these stations.)

a) Predict and write a balanced chemical equation for *each* reaction performed in the lab. For all single replacement and precipitate-forming double replacement reactions *you must* *provide the net ionic equation*. All reactants and products must have states indicated. All ions must have charges!

Reaction	Balanced Chemical Equation
1	
2	
Reaction	Balanced Chemical Equation
3	
4	
5	
6	
7	
8	

9	

Lesson 4.05 Student Notes:

4.06: Reaction Prediction Practice / The Eight Solution Problem Lab

1. (3.B) Sketch a flowchart that you could use to help you predict and write chemical equations based on a written description.

(3.B, 7.E, 8.E, 8.F, 10.B) Predict and write a balanced chemical equation for each reaction. For all single replacement and precipitate-forming double replacement reactions *you must provide the net ionic equation*. All reactants and products must have state signifiers indicated, such as (s), (aq), (l), or (g). All ions must have correct charges!

2. Zinc carbonate powder is heated strongly in a Bunsen burner flame.

3. A cell uses glucose, $C_6H_{12}O_6$, to react with oxygen to generate energy.

4. Solutions of potassium chloride and mercury (I) nitrate are mixed together.

5. Chlorine gas bubbles through aqueous sodium iodide.

6. Powdered sulfur burns in oxygen gas. (Sulfur will have a +4 charge for this one.)

7. Silver nitrate and sodium carbonate solutions are mixed together.

8. A lump of lithium metal is added to water in a test tube.

9. Aluminum foil and fluorine gas are reacted together.

10. Sulfur dioxide and magnesium oxide powder are added together.

11. Zinc is added to a beaker containing copper (II) nitrate solution.

12. Sodium chlorate powder decomposes over a strongly burning Bunsen burner flame.

13. Sodium phosphate and iron (II) sulfate solutions are mixed together.

14. Sulfur dioxide gas mixes with water.

If your teacher wishes, and if you have time, you may be able to complete the following lab.

The Eight Solution Problem Lab

Problem: In this lab, you will react pairs of aqueous solutions to verify if a double replacement reaction has occurred by forming a precipitate. Not all reactions will occur. For reactions that did occur, you will write balanced molecular, total ionic and net ionic equations while identifying the two spectator ions. Once you have some recorded observations, you will be able to identify eight unknown solutions by comparing observations from the known reactions to observations from the reactions involving the unknowns.

Materials per group:

- piece of transparency acetate (1 per group)
- white paper—to add labels & look for precipitate
- known solutions
- unknown, coded solutions

Solutions:

Mercury (I) nitrate
Lead (II) nitrate
Barium nitrate
Silver nitrate
Potassium iodide
Sodium hydroxide
Copper (II) chloride
Iron (III) nitrate

Your unknowns are guaranteed to be within this group of solutions.

Safety:

Lead nitrate and mercury (I) nitrate are toxic. Sodium hydroxide is corrosive. You will not know which solution is your unknown, so assume EVERYTHING is corrosive and toxic. **Contamination is your enemy! Keep everything VERY clean and do not mix up pipettes.**

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

Procedure:

PART 1

1. Inspect your transparency piece. If it appears cloudy, clean it.

2. Draw an 8 x 8 grid on a piece of blank paper. You will get instructions on how to head the rows and columns. Place an X in the appropriate squares of your Knowns Observation Table to show where you do not need to mix solutions.

3. Place the transparency over the grid.

4. Add the indicated solutions, in pairs, on the transparency in each square of the grid. You will be using only 2 drops of each solution for one mixture, for a total of 4 drops for each pair.

NOTE: These materials will be used by the next class so your work cannot be saved. Make all the mixtures and THEN take observations. Sometimes precipitates only form if they are left standing for a few minutes. It may not happen right away – be patient.

5. Make excellent, detailed observations in your "Knowns" data table.

If a reaction does occur, provide further detail. Was there a color change? From what-to-what color? Was a solid formed and what color is it? Is the color bright or pale? Does the solid appear dense, flocculent (clumped) or flaky? Does the solid sink or float or is it suspended in solution? **Do NOT use similes like "looks like milk" or "murky" or "foggy" or "cloudy".**

If no visible reaction occurs, note with the abbreviation "NVR".

6. Take a picture of your completed knowns grid.

7. Dispose of the mixtures by wiping the plastic with paper towel. Rinse if necessary.

PART 2

8. On another piece of paper, draw a new 8 x 8 grid. Along the top of the grid, write the unknown letters, one letter per square. Along the side of the grid, write the formulas of the unknown solutions, one formula per square.

9. Repeat Part 1 for each unknown mixed with each known and take a picture of your completed unknowns x knowns grid.

10. Clean up as you did in Part 1.

Observations: PART 1: Knowns x Knowns Grid

	$Hg_2(NO_3)_2$	NaI	$Ba(NO_3)_2$	$Fe(NO_3)_3$	$AgNO_3$	$Pb(NO_3)_2$	CuCl	NaOH
$Hg_2(NO_3)_2$								
NaI								
$Ba(NO_3)_2$								
$Fe(NO_3)_3$								
$AgNO_3$								
$Pb(NO_3)_2$								
CuCl								
NaOH								

PART 2: Unknowns x Knowns Grid (Write unknown code letters in the first column.)

Unknown	**$Hg_2(NO_3)_2$**	**NaI**	**$Ba(NO_3)_2$**	**$Fe(NO_3)_3$**	**$AgNO_3$**	**$Pb(NO_3)_2$**	**CuCl**	**NaOH**

Analysis:

a) Write a balanced net ionic equation for the reaction of *each* unknown with a known solution that formed precipitate. For example, if NaOH forms a precipitate with Unknown "A" and you think Unknown "A" is copper (II) nitrate, write the three reactions with NaOH and $Cu(NO_3)_2$ as the reactants.

All ions must have charges and all equations must include states and the correct coefficients!

b) Write a conclusion for identifying one of the unknowns (your choice). Identify the unknown, then provide written clarification as to why you believe your unknown to be a particular solution based on how the unknown reacted (or didn't react) with the test solutions.

For example, "I believe unknown Z is hydrochloric acid because it reacted with nitric acid to produce a white solid, reacted with silver nitrate to produce a black solution, and did not react with the other substances. Unknown Z produced identical observations to hydrochloric acid from the Knowns table, therefore unknown Z is hydrochloric acid."

Lesson 4.06 Student Notes:

Unit 5: Moles and Stoichiometry

5.01: Mole Concepts and Molar Mass

1. (4.D) Name the representative particle for each sample of matter. You may need to look up which elements are metals and which are nonmetals. (The Dynamic Periodic Table at https://www.ptable.com should help.)

   Ca NO_2 O_2 AgCl Fe^{2+} H_2SO_4 $Zn(NO_3)_2$

2. (8.A, 8.B) Following the previous example (for $C_6H_{12}O_6$ - glucose) count the number of atoms of each element in 5 formula units of copper (II) phosphate$_{2\ (s)}$, 5 dozen formula units of copper (II) phosphate, and 5 moles of copper (II) phosphate.

3. (8.A, 8.B) Calculate the number of each type of representative particle.

3 formula units sodium oxide:	7 moles calcium nitrate:	4 moles phosphoric acid:
_________ Na^+ ions	_________ moles Ca^{2+} ions	________ moles H^+ ions
_________ O^{2-} ions	_________ moles N atoms	________ moles P atoms
_________ ions in total	_________ moles O atoms	________ moles O atoms

4. (2.G, 8.A) Calculate the molar mass of each substance. Show all your work.

zinc chloride

sodium thiosulfate

lead (II) nitrate

Iron (II) nitrate

Zinc Sulfate

stannic hydrophosphate

5. (2.G, 8.A) You have an unknown sample with a mass of 2.05 g that contains 0.007164 moles. Given that the sample is a compound from #4, calculate the molar mass and identify the unknown sample.

Challenge Question:

6. (2.G, 8.A, 8.B, 3.A) A sample of cyclohexene (C_6H_{10}) is reacted with a sample of pure bromine to create a new compound with the formula $C_6H_{10}Br_x$.

a.) Calculate the molar mass of cyclohexene.

b.) The students used mass spectrometry (*more about this later*) to determine the molar mass of the newly created compound to be 321.858 g/mol. Determine the number of bromines attached to the newly created compound and write the complete chemical formula.

c.) If the students started with 1.84 moles of cyclohexene, calculate the number of moles of bromine used in the reaction to create the sample of the new compound. Assume every molecule of cyclohexene reacted.

Quick Lab: Identify the Unknown Compounds (3.A, 8.A, 8.B)

Examine six small bagged and coded samples of different compounds. Collect data to determine the identity of each compound. Your teacher will give you a list of likely compounds, by formula. The number of moles will be marked on each bag. You have access to an electronic balance.

Data:

Calculations:

Lesson 5.01 Student Notes:

5.02: Mole Conversions

1. (2.G, 8.B) How many atoms are in 3.2 moles of zinc?

2. (2.G, 8.B) How many moles are 4.50×10^{28} molecules of water?

3. (2.G, 8.B) How many nitrogen atoms are in 5.75×10^{24} molecules of dinitrogen pentoxide? How many oxygen atoms are there?

4. (2.G, 8.B) There are 7.44×10^{25} oxygen atoms in a sample of pure magnesium nitrate. How many formula units of magnesium nitrate are in the sample?

5. (2.G, 8.A) Calculate the mass of 0.750 moles of sodium carbonate decahydrate.

6. (2.G, 8.A) Calculate the number of moles of water in a 80.0 mL sample. (Hint: Just like Avogadro's number connects the mole to representative particles; density connects mass to volume.)

7. (2.G, 8.B) How many H^+ ions are there in a 38.2 kg sample of hydrochloric acid? (Show all units, steps and conversions!)

8. (2.G, 8.B) If a sample of ammonium carbonate contains 3.077×10^{21} formula units, what is the mass of the sample in mg?

9. (2.G, 8.B) A sample of an unknown compound is known to contain 5.00×10^{24} nitrogen (N) atoms. If the compound is pure copper (II) nitrate, what is the mass of the sample?

For mixtures, calculate the number of moles of atoms for each component and then add them together. You must do this because the different components will each have different molar masses.

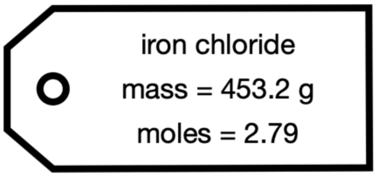
10. (2.G, 8.B) How many potassium atoms are in a 27.9 cg mixture of 60.0% potassium sulfate and 40.% calcium sulfate?

11. (2.G, 8.B) What is the mass of *just* the oxygen atoms in the 27.9 cg mixture from the previous question? (Use 15.999 g/mol as the molar mass of the oxygen atoms.)

12. (8.B) Calculate the number of sodium atoms in a sample of 54.67 g of naproxen (Aleve©) $C_{14}H_{13}NaO_3$.

Challenge Question:

13. (8.A) A student has obtained a sample of iron chloride ($FeCl_x$) from the stock room. Use the bottle label to determine the molar mass and formula of the compound in the bottle.



iron chloride
mass = 453.2 g
moles = 2.79

The Great Salt Challenge (2.E, 2.F, 2.G, 8.B)

Purpose*:* You will be given a specific number of formula units of salt. You must calculate to determine the mass of the salt to give to your teacher. Table salt is sodium chloride.

Your teacher will give you the following information: The number of formula units of salt you need to obtain is: ______________________________

Procedure*:* Design a procedure so that you can calculate the amount of salt you need to measure out. Write it out in a list of numbered steps so a stranger can follow your lab and get your results.

Calculations:

Use dimensional analysis to calculate the mass of chemical you need to mass out.

What mass of salt do you need to obtain? ___________________________

Mass out the sample of salt you need and present it to your teacher in a weighing boat.
You will need to provide the mass of an empty weighing boat in your data.

Data:

Mass of empty weighing boat (g)	
Mass of weighing boat + salt (g)	

Analysis:

a) Use dimensional analysis to calculate the number of sodium ions in your sample.

b) Calculate the mass of just the sodium atoms in your sample of salt. Use the molar mass of sodium in your dimensional analysis.

c) Is the number of sodium ions in your sample greater than, equal to, or less than the number of chlorine ions in your sample? Provide justification for your answer.

d) Is the *mass* of sodium ions in your sample greater than, equal to, or less than the *mass* of chlorine ions in your sample? Provide justification for your answer in the form of calculations and/or a discussion of the concept of a mole.

Lesson 5.02 Student Notes:

5.03: Percent (%) Composition, Empirical and Molecular Formulas

Percent Composition of Gum (2.E, 2.F, 2.G, 2.H, 2.I, 3.B, 8.C)

With your partners, design an experiment that you will perform to determine the % of sugar in a piece of gum. You have available: a piece of gum, and access to a balance. You do not need to know any nutritional information (promise!).

Write your steps here:

Write your calculations here and state the % sugar in your piece of gum here. Consider and use the correct number of significant figures in your final % value. Record your value on the board if your teacher asks.

What sources of error contribute to any differences in the data between groups? Explain.

1. (2.G, 8.C) Calculate the percent composition by mass of each element in magnesium nitrate.

2. (2.G, 3.B, 8.C) Which is a better source of oxygen, silver nitrate or water? Use calculations to support your conclusion.

3. (2.G, 8.C, 8.D) Calculate the empirical formula of a compound which contains 40.04% calcium, 12.00% carbon, and 47.96% oxygen.

4. **(**2.G, 8.C, 8.D) Calculate the empirical formula for a compound which contains 26.8% tin, 16.0% chlorine and 57.2% iodine.

5. (2.G, 8.C, 8.D) A compound is 12.7% aluminum, 19.7% nitrate, and 67.6% oxygen. Determine its empirical formula.

6. (2.G, 8.C, 8.D) Boron hydrides, compounds containing only boron and hydrogen, form a large class of compounds. One consists of 78.14% boron and the rest is hydrogen. The molar mass is 27.7 g/mol. What are the empirical and molecular formulas of this compound?

7. (2.G, 8.C, 8.D) A compound with the following composition has a molar mass of 60.10 g/mol. It is 39.97% carbon, 13.41% hydrogen, and 46.62% nitrogen. Find the molecular formula.

8. (2.G, 8.C, 8.D) Estradiol is a hormone that causes maturation and maintenance of the female reproductive system. Elemental analysis of estradiol gives the following percent composition: carbon 79.37%, hydrogen 8.88%, oxygen 11.75%. The molar mass of estradiol is 272.37 g/mol. Determine the empirical and molecular formula of estradiol.

9. (8.C, 8.D) NutraSweet© is a sweetener with a molar mass of 294.30 g/mol. It is 57.14% carbon, 6.16% hydrogen, 9.52% nitrogen, and 27.18% oxygen. Calculate the empirical and molecular formula of this sweetener.

10. (2.G, 8.D) Caffeine is 49.47% carbon, 5.201% hydrogen, 28.84% nitrogen, and the rest are oxygen. If the molar mass of caffeine is 194.2 g/mol, calculate the empirical and molecular formula of caffeine.

11. (8.A, 8.C) A metal, M, forms a chloride with the formula MCl_3. If the compound contains 65.57% chlorine by mass, what is the identity of the metal? (Hint: determine the molar mass of M.)

Challenge Question

12. (2.G, 3.B, 8.B, 8.C, 8.D) A method of determining the empirical formula of a hydrocarbon (a compound only containing C's and H's) is through combustion analysis (i.e., burning it). Burning a hydrocarbon always produces carbon dioxide and water. You burn an unknown hydrocarbon, $C_xH_{y\,(g)}$, in oxygen gas and mass the products. The reaction produced 110 g carbon dioxide and 67.5 g water.

a.) Calculate the number of moles of carbon dioxide and water produced.

b.) Determine the number of moles of C and H in the original compound (*Hint*: Atoms can't be created or destroyed! Thus, the moles of carbon atoms we end with must be how many we started with.)

c.) Determine the empirical formula (*Hint*: See Step 3 of percent composition example) d.) Using mass spectroscopy, the molar mass of the molecular formula was determined to be 45.105 g/mol. Determine the molecular formula of the unknown hydrocarbon.

Lesson 5.03 Student Notes:

5.04: Hydrate Calculations

1. Calculate the molar mass of each substance. Show all your work.

    sodium carbonate decahydrate
    calcium chloride dihydrate
    barium hydroxide octahydrate

    sodium tetraborate decahydrate
    sodium sulfate decahydrate
    magnessium sulfate heptahydrate

2. (3.B) Draw a particle diagram representing one formula unit of magnesium carbonate pentahydrate.

3. (2.G, 3.B, 8.A) A 7.028 g sample of a sodium sulfate hydrate is heated over a Bunsen burner flame in a crucible. After all the water is driven off, 3.100 g of the anhydrous compound is left over. Determine the formula of the hydrate using calculations.

4. (2.G, 3.B, 8.A) A cobalt (II) chloride hydrate is heated in a crucible to remove all the water, and the data is given in the table below.

Mass of crucible (g)	12.090 g
Mass of crucible and hydrated salt (g)	12.701 g
Mass of crucible and anhydrous salt (g)	12.424 g

Using the data and calculations, determine the formula of the hydrate.

5. (2.G, 8.C) Find the mass percent of water using the hydrate formula you determined in practice problem #2.

6. (2.G, 8.C) Find the mass percent of water using the hydrate formula you found in practice problem #3.

7. (2.G, 3.B, 8.B, 8.C) A sample of copper (II) chloride hydrate has a mass of 3.41 g. After all the water is driven off, the anhydrous compound weighs 2.69 g.

a) Calculate to determine the IUPAC formula for this hydrate.

b) Calculate the number of chloride ions in the anhydride.

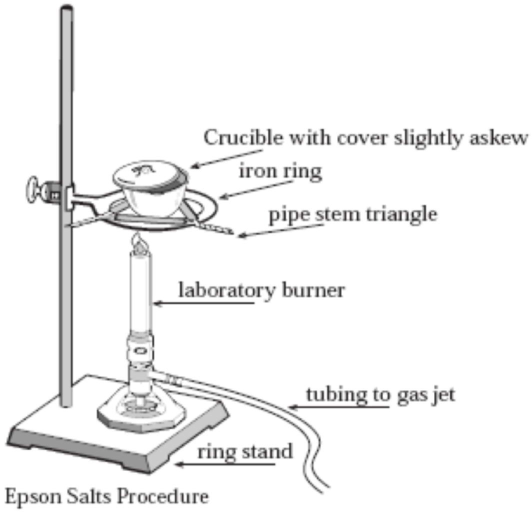
c.) Calculate the % composition of oxygen and copper in the copper (II) chloride hydrate.

8. (3.B) Determine the chemical formula of a hydrate that is 76.9% calcium sulfite and 23.1% water. Provide calculations to support your answer.

Lesson 5.04 Student Notes:

5.05: Lab: Waters of Hydration in Epsom Salt

Problem*:* Determine the waters of hydration for Epsom salt, a magnesium sulfate hydrate.



Background: The waters of hydration will be driven off by heating over a low Bunsen burner flame. When the salt has stopped changing mass (i.e., becomes anhydrous), you can determine the mass of the evaporated water and the mass of the anhydrous salt. The ratio of the moles of water to the moles of anhydrous salt will give you the number of waters of hydration, allowing you to write a formula for the hydrate.

Gentle heating is required to ensure the evaporating dish does not crack and that you remove *just* the water. You are not decomposing the magnesium sulfate![6]

Procedure*:*

1. Wash and dry the evaporating dish.

2. Set up your equipment as in the image on the previous page, except you are using an evaporating dish, not a crucible with cover.

3. Light the burner and gently heat the evaporating dish for a couple of minutes until any water has evaporated.

4. Turn off the burner and cool until it is comfortable to the touch. Record the mass of the dry evaporating dish.

5. Add about 5 g of Epsom salt to the evaporating dish. Record the mass of the evaporating dish plus Epsom salt.

6. Place the evaporating dish back in the pipe stem triangle, light the burner again, and gently heat until all the water has been released from the salt. You may see “bubbling” or “boiling”.

---

[6] Figure taken from *Epsom Salts Procedure*. whs.rocklinusd.org/documents/Science/Epsom_salt_lab.pdf.

7. Turn off the burner and allow the evaporating dish and dehydrated salt to cool completely. This may take several minutes. Record the mass of the evaporating dish and dehydrated salt.

8. Relight the burner and gently reheat the dish for a few minutes. Cool and weigh again. If the two masses (evaporating dish plus dehydrated salt, measured twice) are VERY close, you do NOT have to heat again. If they are not… repeat step 8.

You must perform at least ***two*** trials and record the data and calculations for each trial.

The dehydrated Epsom salt may be placed in the trash after the dish has completely cooled.

Claim: Your group will write a complete sentence for the claim that states what you believe the formula of hydrated Epsom salt will be, i.e. "The formula of the hydrated Epsom salt is ___________."

Data (2.F):

If your group needed to do more heatings, please extend this data table to include those values.

	Trial 1	Trial 2	Trial 3 (opt.)
Mass of empty evaporating dish (g)			
Mass of dish + Epsom salt (g)			
Mass of dish + anhydrous salt after 1st heating (g)			
Mass of dish + anhydrous salt after 2nd heating (g)			

Data Analysis:

Show ALL of your work for all trials. Use sig. fig. rules and a unit after every answer.

1. (2.F, 2.G) Calculate the mass of the evaporated water. (Use the lowest mass of dish + anhydrous salt.)

2. (2.G, 8.A) Calculate the number of moles of the evaporated water.

3. (2.F, 2.G) Calculate the mass of the anhydrous salt. (Use the lowest mass of dish + anhydrous salt.)

4. (2.G, 8.A) After you determine the chemical formula for magnesium sulfate, calculate the number of moles of the anhydrous salt.

5. (8.A) Divide the number of moles of evaporated water (*answer from #2*) by the number of moles of anhydrous salt. This ratio is your waters of hydration. Round if necessary.

Discussion (2.H, 2.I): Explain how specific parts of the evidence help support your claim that the formula of Epsom salt is what you say it is. Please refer to scientific background from the course notes as you relate the claim to the data.

5.06: Stoichiometry: The Basics

1. (8.G) Write all the mole ratios for $2\ C_2H_{6\ (g)}\ +\ 7\ O_{2\ (g)} \longrightarrow 4\ CO_{2\ (g)}\ +\ 6\ H_2O_{\ (g)}$.
   (Hint: There are six of them.)

2. Write a balanced chemical equation for the reaction of aqueous sulfuric acid with lithium hydroxide solution. Use it to solve the following stoichiometric problems. Consult Units 3 and/or 4 for help if necessary.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G) How many moles of sulfuric acid can react with excess lithium hydroxide to produce 4.3 moles of water?

b) (2.G, 8.G) How many moles of lithium hydroxide are required to react with 1.23 moles of sulfuric acid?

If you're not provided with a balanced chemical equation, you will need to write one, including the state signifiers, BEFORE you start stoichiometry.

3. (2.G, 8.G) Methane gas, $CH_{4\ (g)}$, combusts in excess oxygen.

Write the balanced chemical equation here, including the state signifiers.

a) Calculate the mass of oxygen needed to react with 13.6 g of methane.

4. Bromine gas reacts with aqueous sodium iodide, including state signifiers.

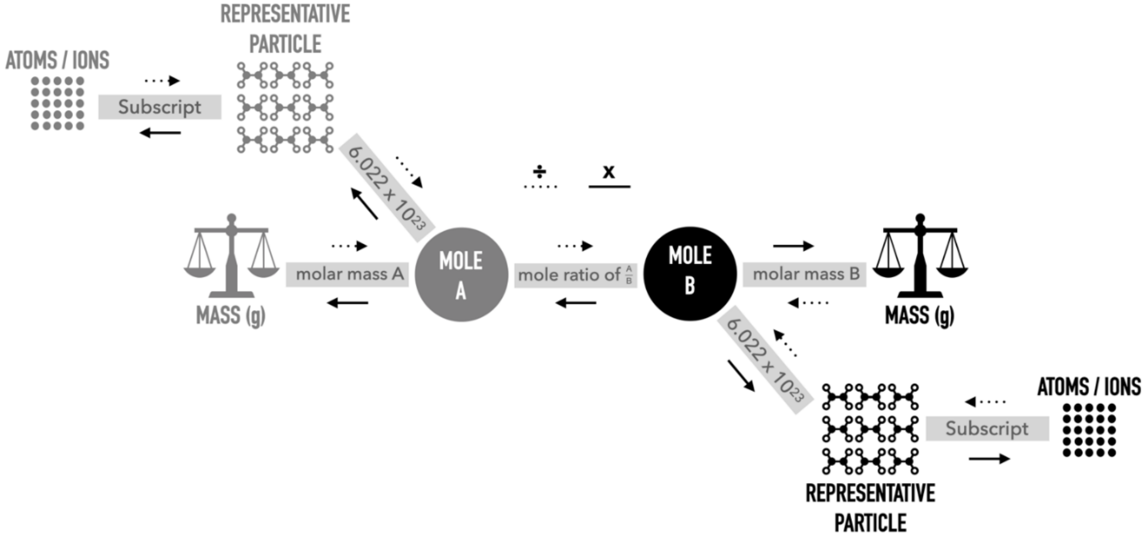
Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G) How many moles of I_2 are produced when 3.00 mol of bromine completely reacts with excess sodium iodide?

b) (2.G, 8.G) How many mg of sodium iodide will completely react with 10.0 mol of bromine?

c) (2.G, 8.G) How many grams of iodine solid could be produced when 50.0 g of bromine completely reacts with excess sodium iodide?

For the next two problems, we will combine stoichiometry with what you learned in 5.01. You can also think of it as the following larger mole map.



d) (2.G, 8.B, 8.G) How many formula units of sodium bromide could be produced from 0.172 mol of bromine?

e) (2.G, 8.B, 8.G) How many molecules of bromine gas are required to create 7.89 x 10^{34} molecules of sodium bromide?

5. (2.G, 3.B, 8.B) A student was asked to determine the mass in grams of hydrogen gas (H_2) that would be needed in the synthesis of 10 mg of ammonia (NH_3).

They created the following set of "train tracks". Identify the error in the student's train tracks and explain how to correct it.

	Step 1	**Step 2**	**Step 3**	**Step 4**
10 mg NH_3	1 g NH_3	1 mole of NH_3	3 moles of H_2	1.008 g of H_2
	1000 mg NH_3	17.031 g of NH_3	2 mole of NH_3	1 mole of H_2

6. (2.G, 8.B, 8.G) A reaction occurs between iron (III) chlorate and calcium hydroxide solutions. How many kilograms of each product can be produced given that 2.56 g of iron (III) chlorate was reacted with excess calcium hydroxide?

Lesson 5.06 Student Notes:

5.07: Stoichiometry: Theoretical and Actual Yield

1. (2.G, 8.G) You calculate the theoretical yield of an experiment to be 4.50 g. When you perform the experiment the actual yield you obtained was 3.97 g. What is the percent yield of your experiment?

a) (2.G, 8.G) Using the data from #1, does this represent a loss of product or do you have too "much" product?

b) What specific systematic errors occurred that would explain this discrepancy?

2. Potassium chlorate decomposes into solid potassium chloride and oxygen gas.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G) What is the theoretical yield of potassium chloride if 12.00 moles of potassium chlorate decomposed?

b) (2.G, 8.G) In an experiment, 12.00 moles of potassium chlorate decomposed to yield 800. g of potassium chloride. What is the percent yield of potassium chloride for the experiment?

c) (2.G, 8.B, 8.G) How many potassium ions are present in the actual yield of potassium chloride?

3. Using the following reaction, camels store the fat tristearin ($C_{57}H_{110}O_6$) in the hump. As well as being a source of energy, the fat is a source of water.

$$2\ C_{57}H_{110}O_{6\ (s)} \quad + \quad 163\ O_{2\ (g)} \quad \text{-->} \quad 110\ H_2O_{\ (g)} \quad + \quad 114\ CO_{2\ (g)}$$

a) (2.G, 8.G) What is the theoretical yield of carbon dioxide in grams if 8.63 kg of tristearin is burned during a lab experiment?

b) (2.G, 8.G) The actual yield of the experiment above is 18.7 kg. What is the percent yield?

4. Ammonia gas reacts with carbon dioxide gas to form liquid water and solid urea, $CN_2OH_{4\ (s).}$

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.B, 8.G) 3.49 x 10^{21} ammonia molecules react completely with excess carbon dioxide. What is the theoretical yield of urea in cg?

b) (2.G, 8.G) If 13.74 cg of urea are produced from the experiment above, what is the percent yield?

Challenge Question:

5. (2.G, 3.B, 8.A, 8.G) For this problem, you will be investigating the creation of propene (C_3H_6) through a process called steam cracking. Watch the following video to learn more about it: https://bit.ly/2YBoNQG.

   Propene is used to create so many by-products such as paints, electronics, auto parts, apparel, carpet, surface coatings, adhesives, etc[7]. An example of this cracking is breaking propane (C_3H_8) up into propene and hydrogen gas, as shown below.

$$C_3H_{8\,(g)} \rightarrow C_3H_{6\,(g)} + H_{2\,(g)}$$

a.) Calculate how many grams of propene can be produced from 612 moles of propane.

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[7] American Chemistry Council, *Guide to the Business of Chemistry.* 2019.

b.) The cracking of propane does not produce just propene. Typically, about 17% of the products is propene, 42% in ethylene, 28% is methane and the rest are a mixture of other compounds[8]. Using this product breakdown, calculate how many moles of propane you would need to crack in order to acquire 28.6 kg of propene.

c.) It costs \$2.48/g [9] for propane. If you were to produce 28.6 kg of propene, which sells for \$1.65/g, calculate how much money you would lose or gain in the process considering the other by-products as waste.

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[8] Faisca, Nuno. From Crude Oil to Chemicals: Back to the Fundamentals. Nexant, 2018.

[9] Sigma Aldrich, 2020.

Lesson 5.07 Student Notes:

5.08: Stoichiometry: Limiting and Excess

Use the following chemical equation, to draw a particle diagram for each scenario. Put each reactant in their own box, and all of the products (and the leftover excess reactant) into the third box.

1. Sodium metal reacts with water.

Write the balanced chemical equation here, including the state signifiers.

a) (8.H) 6 atoms of Na and 4 atoms of H_2O

b) (8.H) 7 atoms of Na and 6 atoms of H_2O

2. Propane, $C_3H_{8\ (g)}$, undergoes combustion.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G, 8.H) When 0.56 moles of propane react with 0.64 moles of oxygen, how many moles of carbon dioxide are formed?

b) (2.G, 8.G, 8.H) For the reactant in excess, how many moles are left over at the end of the reaction?

c) (2.G, 8.B, 8.G, 8.H) How many *molecules* of water were produced during the reaction?

3. One step in the extraction of copper from copper ores is “roasting”. The roasting of copper ore, $CuFeS_{2\ (s)}$, entails combusting the ore in oxygen gas to form iron (II) oxide, copper (II) sulfide, and sulfur dioxide gas.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G, 8.H) If 20.0 g of each reactant is mixed together, calculate the theoretical yield of copper (II) sulfide and the theoretical yield of sulfur dioxide.

b) (2.G, 8.G, 8.H) What mass of excess reactant is left over from the previous question?

4. The reaction between potassium chlorate and red phosphorus ($P_{4\ (s)}$) takes place when you strike a match on a matchbox. It produces tetraphosphorus decaoxide solid and potassium chloride solid.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G, 8.H) What is the theoretical yield of tetraphosphorus decaoxide when 2.50 g of potassium chlorate and 5.00 g of red phosphorus react?

b) (2.G, 8.G, 8.H) What is the mass of excess reactant left over?

c) (2.G, 8.G, 8.H) How much more of the limiting reactant would you need to react with the excess reactant?

5. Using the following equation, reusable booster rockets of the U.S. space shuttles employ a mixture of aluminum and ammonium perchlorate for fuel.

$$Al_{(s)} \quad + \quad NH_4ClO_{4\,(s)} \quad \text{-->} \quad Al_2O_{3\,(s)} \quad + \quad AlCl_{3\,(s)} \quad + \quad NO_{(g)} \quad + \quad H_2O_{(g)}$$

a) (2.G, 8.G, 8.H) If 25.0 kg of $Al_{(s)}$ react with 32.0 kg of NH_4ClO_4, how many kilograms of $Al_2O_{3\,(s)}$ can be made?

b) (2.G, 8.G, 8.H) If 27.0 kg of $Al_2O_{3\ (s)}$ are actually produced, what is the percent yield?

Challenge:

6. (3.B, 3.E, 8.G, 8.H) It is your turn to explore chemistry in a new light. There is a vast sea of options when it comes to chemistry and the way it interacts with all disciplines. Take the time to think about chemistry as it relates to your career goals. This might be directly like working as pharmaceutical chemist or understanding the chemistry happening with in your own body. It might also be a bit more abstract like designing the perfect drink, thinking of new ways to paint structures, or creating the next microprocessor for your phone.

a. Explain how chemistry plays a part in your future career.

b) Find one source to learn more about it. Provide 3 takeaways and refer to the source in your explanations.

c) Find one chemical reaction associated with your career in chemistry. If you can't find one, any will work.

d) Using the reaction you found in c), write your own limiting and excess stoichiometric problem and answer it. Now it's your turn to challenge me…

Lesson 5.08 Student Notes:

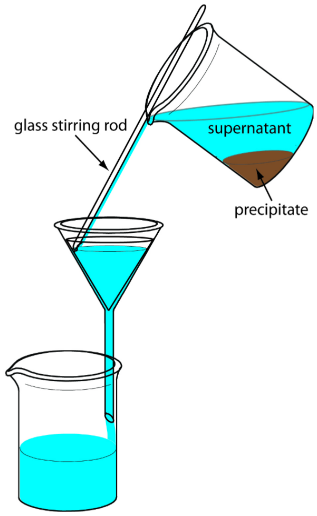
5.09: Gravimetric Analysis

Gravimetric analysis is a lab technique where you first predict the mass of a product you hope to make through a chemical change, physically manufacture and recover that product, and then compare the mass you predicted to the mass you really did make. The comparison at the end is important because it serves as a check on the number of errors (systematic or random) that occurred.

The mass you predict is your theoretical yield of product, and you would predict this by doing a mass-to-mass stoichiometry problem. The mass you really make, through experimentation, is your actual yield. You can then calculate the percent yield. The higher the percent yield, the fewer and/or more insignificant errors were made. A lower percent yield means loss of reactants, loss of product, or both. A higher percent yield means there is excess "product"; you may have unwanted products or unwanted substances mixed in with your product, which is adding to your actual yield. Many times, an unwanted "side reaction" (some other chemical reaction, other than the one you are using or studying) is yielding products that add to the product you set out to create. In industry, it is incredibly common to obtain low percent yields. Improving and streamlining procedures to promote efficiency will improve percent yield and thus increase revenue.
The procedure usually goes as follows: [10]

Procedure:

- Weigh a piece of dry, clean filter paper.
- Perform a chemical reaction that will form the product of interest.
- Separate the solid – or the **precipitate** – from unused reactants and waste product. (This "waste" liquid is called the **supernatant**.)
- To separate them, we filter the precipitate through filter paper and a funnel so the supernatant passes through the funnel to a beaker below. See the graphic on the left. This helps prevent product loss down the funnel.



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[10] Libretexts. "8.2: Precipitation Gravimetry." *Chemistry LibreTexts*, Libretexts, 24 Dec. 2016, chem.libretexts.org/LibreTexts/Northeastern/08%3A_Gravimetric_Methods/8.2%3A_Precipitation_Gravimetry. Licensed under CC BY NC SA 3.0: https://creativecommons.org/licenses/by-nc-sa/3.0/us/legalcode.

- Wash the precipitate several times, with distilled water, to ensure all the waste chemicals have completely removed before drying. Simply add distilled water to the precipitate and filter paper gently, so as not to lose product.
- Once you have finished filtering, dry the precipitate thoroughly. Weighing the precipitate while wet adds unwanted mass. Weigh the dried precipitate and filter paper. Ideally, you would weigh it several times, then record the lowest mass reached.

Here are some practice problems that will help you get a grasp on this important technique. Remember that these are all just applications of mass-to-mass stoichiometry; it is just some of the language that is changing.

1. During a gravimetric analysis experiment, you add 2.00 g of solid calcium chloride to an aqueous solution of sodium carbonate. After the calcium chloride has dissolved, a precipitation reaction can take place. The precipitate is filtered, washed, dried, and weighed.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G) Calculate the theoretical yield of the precipitate. (Assume the calcium chloride is limiting.)

b) (2.G, 8.G) Using the data given below, calculate the actual yield as obtained through an experiment.

Mass of filter paper	11.2083 g
Mass of filter paper and precipitate	12.5921 g

c) (2.G, 3.B) Using the actual yield, calculate the mass of calcium chloride that truly reacted. Was it really 2.00 g? From this information, infer whether the calcium chloride was limiting or in excess and explain your answer.

d) (2.G, 8.G) Calculate the percent yield of the reaction.

Gravimetric Analysis Lab

Purpose: Use stoichiometry and lab skills in the process of **gravimetric analysis** to produce solid calcium carbonate.

Materials:

Calcium chloride, $CaCl_{2\ (s)}$
Sodium carbonate, $Na_2CO_{3\ (s)}$
Weighing boats
Stirring rods
Water
3 beakers
Filter paper and funnel
Graduated cylinder

Procedure:

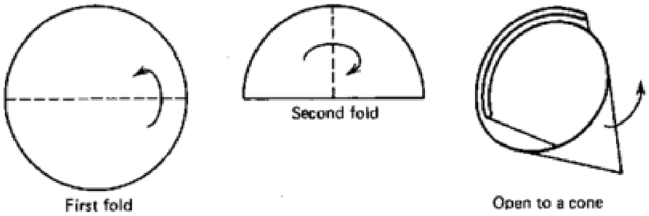
1. Your teacher will provide the mass of calcium chloride and sodium carbonate you will use. Write these masses here: _______________ (calcium chloride) and ________________ (sodium carbonate).

2. Obtain these masses in the weighing boats provided. Note: you do not have to get the mass exactly correct. Get them to within 0.1 g of the masses your teacher gave you. Record these masses in the data table below.

3. Dissolve each solid in separate beakers using a small volume of water (~15 mL or less). Label the beakers if they are not already labelled. Stir well to make sure all the solids dissolve completely.

4. Pour both aqueous solutions of solid into the third beaker so they mix and react. Clean your stirring rod, then stir the mixture in the third beaker to ensure complete reaction. Allow the mixture to settle. The solid calcium carbonate will be visible at the bottom.

5. Weigh and record the mass of one piece of filter paper. Fold the filter paper using the graphic below. Once you have the cone made, place it in the funnel. Someone may have to hold the cone in the funnel the whole time.



[11]

6. Filter the supernatant/precipitate mixture *slowly* using the method shown on pg. 153. It is essential that you get as much of the precipitate onto the filter paper.

7. Rinse the (third) beaker and the stirring rod with small amounts of water. Put this liquid down the stirring rod to catch the precipitate. You will also (probably) need to refilter the supernatant if it is cloudy.

8. Carefully unfold the filter paper and lay it on paper towel to dry.

9. Weigh the dried filter paper and precipitate using the same electronic balance as before. Record this mass in your data table.

Data (2.F):

Mass of weighing boat #1 (g)	
Mass of weighing boat #1 + calcium chloride (g)	
- Calculated mass of calcium chloride (g)	
Mass of weighing boat #2 (g)	

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[11] Hankova, Sonia. "Separating Insoluble Substances." http://sonhank.com/wp-content/uploads/2015/10/Mixtures-FiltrationExperiment.pdf. 2015.

Mass of w.boat #2 + sodium carbonate (s) (g)	
- Calculated mass of sodium carbonate (g)	
Mass of filter paper before folding (g)	
Mass of dry filter paper + calcium carbonate precipitate (g)	
- Calculated mass of precipitate on paper (g)	

Calculations: You MUST show all work! Use sig fig rules to round your final answer and place the appropriate unit after it.

a) Write and balance a chemical equation for the reaction of calcium chloride with sodium carbonate solution.

b) (2.G, 2.H, 8.G) Using the calculated masses of reactants from the data table above, calculate the theoretical yield of solid calcium carbonate. Hint: you will first need to determine which is the limiting reactant.

c) (2.G, 8.G) Calculate the percent yield of calcium carbonate precipitate. Hint: the "calculated mass of precipitate on paper (g)" is the actual yield.

Analysis: (2.H, 2.I) *Explain* the effect of each event on the percent yield – does it make the percent yield too high, too low, or have no effect? Use lots of detail!

a) Students did not refilter their supernatant before drying and weighing their precipitate.

b) Students weighed wet filter paper and precipitate and used that data in their calculations.

c) The electronic balance used by a student group consistently weighed masses 1 gram too low.

d) Students did not weigh the filter paper before using it to filter the supernatant/precipitate mixture.

Lesson 5.09 Student Notes:

Unit 6: Thermochemistry and Kinetics

6.01: Thermodynamics, Heat and $q = mC\Delta T$

1. (11.A, 11.B) A fire is started in a firepit by striking a match and lighting pine needles under some logs. Explain all the energy transfers in this scenario using the terms: exothermic, endothermic, system, surroundings, potential energy, and kinetic energy in your discussion.

2. (11.A, 11.B) When you go into a kitchen and touch a stainless-steel refrigerator handle, it feels colder than a plastic cup you pick up. Explain all the energy transfers using the terms: exothermic, endothermic, system, surroundings, potential energy, and kinetic energy in your discussion. Make sure to explain why the handle feels colder than the cup.

3. (2.G) Use dimensional analysis to convert between the following quantities and units.

a) 37.5 J to calories

b) 0.0500 Calories to J

c) 150 kJ to calories

d) 0.175 calories to kJ

4. (4.B) Are the following quantities intensive or extensive physical properties? Justify your answers.

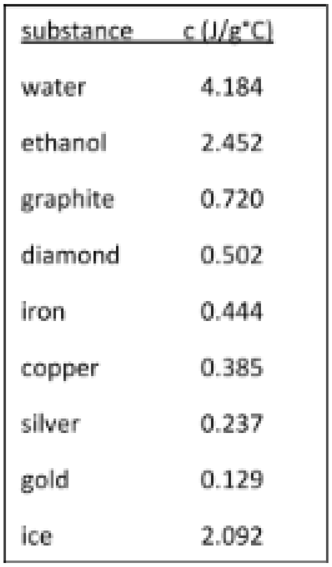
a) Specific heat capacity (C)

b) Energy or heat (q)

c) Mass (m)

d) Change in temperature

Use the chart of specific heat capacities to help you solve all q = mCΔT problems in the course notes. You are required to put a sign (+ or -) on all values for *q*. If energy is lost from the system, *q* will be negative. If the system gains energy, *q* will be positive.

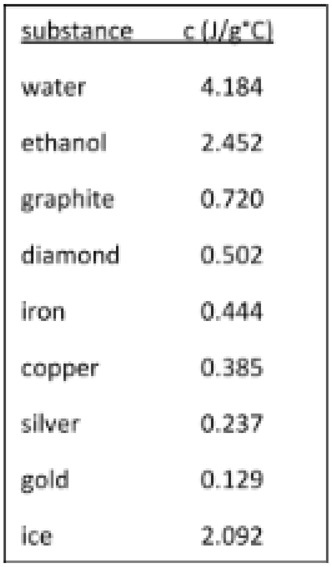


substance	c (J/g°C)
water	4.184
ethanol	2.452
graphite	0.720
diamond	0.502
iron	0.444
copper	0.385
silver	0.237
gold	0.129
ice	2.092

5. (2.G, 11.D) How much heat is produced when 654 g of water temperature raises from 34.5°C to 89.7°C?

6. (2.G, 11.D) How much heat is produced when 654 g of silver temperature decreases from 89.7°C to 34.5°C?

7. (2.G, 11.D) If 7350 J were added to 152 g of ethanol, what *change* in temperature will you observe?



substance	c (J/g°C)
water	4.184
ethanol	2.452
graphite	0.720
diamond	0.502
iron	0.444
copper	0.385
silver	0.237
gold	0.129
ice	2.092

8. (2.G, 11.D) 16.25 g of water at 54.0˚C produces 402.7 J of energy. What will be its final temperature?

9. (2.G, 11.D) 100. g of graphite gains 3000. J of energy to reach a final temperature of 81.7℃. What was its initial temperature?

10. (2.G, 11.D) 697 J are added to a 36.8 g of kerosene and the temp increases from 22.5˚°C to 34.7˚C. Determine kerosene's specific heat.

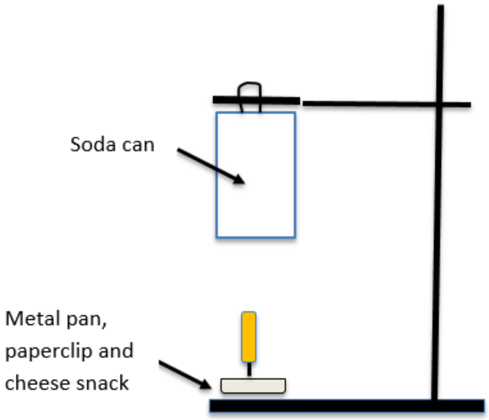
Energy of Food Lab

Problem: What is the heat of combustion of two different types of cheese snack in Joules / g and Calories / g? How many Joules and Calories are truly in a serving of snacks? Are these consistent with the published Calories per serving on the bag?

Materials:

Ring stand, iron ring
Soda can
Water
Thermometer
Stirring rod threaded through soda can
Foil/metal pan
Unwound paperclip
Cheese snacks
Candle lighter or match
100 mL graduated cylinder

Procedure:



1. Move your bags away from the lab tables.
2. Place < 50 mL of water in the empty soda can, using the graduated cylinder. Record the volume of water with all certain digits (the markings) plus one uncertain digit.
3. Weigh one cheese snack, the foil, the metal pan, and the paperclip *together*, and record this mass.
4. Mount the snack on the foil and paper clip.
5. Move the ring clamp so that the fire will *not* touch the bottom of the can but will be close enough to heat the water through the can.
6. Insert the thermometer into the water (not the can, the water!) and record its temperature with all certain digits (the markings) plus one uncertain digit.
7. Set the snack on fire and allow it to burn completely out. You may have to re-light.
8. After it has burned out, record the temperature of the water. Weigh the burned snack, the foil, the metal pan, and the paperclip together.
9. Repeat for one more trial of the same type of snack, using a new snack every time.
10. Repeat steps #3-9 for a second type of cheese snack, if your teacher directs you to.

__Data__: (2.F, 2.G) Use sig fig rules and a unit after each of your entries.

Snack Type: ________________________
Mass of one serving of snack from bag label: ________________
Number of snacks per serving from the bag label: ________________

Measurement	Trial 1	Trial 2	Trial 3
Mass of one snack, foil, metal pan, paperclip before burning			
Mass of one snack, foil, metal pan, paperclip after burning			
Mass of water ** One uncertain digit required ** density of water = 1 g/mL			
Initial temp. ** One uncertain digit required			
Final temp. ** One uncertain digit required			

__Calculations:__ All work and a formula must be shown! Use sig fig rules and a unit after each of your final answers.

(2.G, 11.D) Calculate the heat of combustion and the number of Calories in a serving of your cheese snack.

Calculation	Trial 1	Trial 2	Trial 3
Calculate the mass of snack that burned.			
Calculate ΔT of the water.			

Calculate *q* gained by the water: q = mCΔT. Look up C (H_2O) on pg. 12. This is equal to the *q* lost by the burning snack.			
Calculate the heat of the snack (*q*) *per g* of snack that burned, in J/g. *This value is the heat of combustion in J/g.* Remember: per = divide			
Calculate the heat of snack per g burned in units of Calories / g. *This value is the heat of combustion in Calories/g.*			
Calculate the Calories in one serving of cheese snack.			

Analysis: Write in complete sentences to answer the following questions.

a) (11.A) Describe how the different types of energy in the snack, such as potential, thermal, and kinetic energy, are being transformed to warm the water in the soda can. Make sure to describe what is going on at the particle level.

b) (11.B) How does this experiment demonstrate the zeroth and first laws of thermodynamics? Be specific, using the terms "system", "surroundings", and "universe" in your response.

c) (2.H, 2.I, 3.B) Describe **AND EXPLAIN** two sources of error that cause your calculated Calories/serving to be different from that stated on the bag. Explain what other variable(s) is/are too high or too low that directly causes the calculated Calories/serving to be too high or too low. (Hint: these errors will effect mass of Cheeto, mass of water, or temperature change of water.)

Challenge Problems:

11. (2.G, 11.D) Two containers of water are at 20℃ initially. One contains 25 mL and the other 50 mL. They are each heated with the same source of heat for the same amount of time. If the final temperature of the 25 mL sample is 50℃, what would be the final temperature of the 50 mL sample?

12. (2.G, 11.D) Two containers each have 50 mL of water at 20℃ initially. They are each heated with the same source of heat, but one is heated for 10 minutes and the other for 5 minutes. If the container that was heated for 5 minutes has a final temperature of 30℃, what would be the final temperature of the other sample?

Lesson 6.01 Student Notes:

6.02: Heat and Calorimetry I and II; Specific Heat of Metals Lab

1. (2.G, 11.D) A lump of chromium metal has a mass of 95.3 g and a temperature of 90.5℃. It is placed into a calorimeter with 75.2 mL of water at 20.5℃. After thermal equilibrium has been reached, the final temperature of the water, metal, and calorimeter is 28.6℃. What is the specific heat of chromium?

2. (2.G, 11.D) A 25.0 g sample of zinc is heated, then placed in a calorimeter containing 65.0 g of water at 20.00℃. The temperature of the zinc and water stabilizes to 22.50℃. What was the initial temperature of the zinc metal sample? C of zinc = 0.387 J/(g·℃).

3. (2.G, 11.D) A sample of 50. mL of water at 80.°C is added to 50. mL of water at 20.°C. What would be the final temperature of the total volume of water?

4. (2.G, 11.D) What mass of 54.0°C water must be added to 468 g of 21.0°C water to make the final temperature of both come out to be 29.0°C?

5. (2.G, 11.A, 11.D) 100.0 g of water at 20.0°C are mixed with 200.0 g of copper at 40.0°C.

a) Draw a particle diagram illustrating the heat flow occurring between the water and the copper before thermal equilibrium occurs.

b) Calculate to determine the final temperature.

Challenge Problems:

6. (2.G, 11.D) 25.0 g of glass (C = 0.749 J/g℃) is added to 150. g of water. How much energy is needed to raise the temperature of *both the glass and the water* from 20.0℃ to 50.0℃ in calories?

Lesson 6.02 Student Notes:

6.03: Enthalpy of Phase Changes

1. (4.A, 4.C) Define and draw a particle model that depicts the following phase changes.

Phase Change	Definition	Particle Diagram
Melting		
Freezing		
Evaporating		
Condensation		
Sublimation		

Deposition		

2. (3.B, 4.A, 4.C, 11.A) Write a description of what is happening on the particle level in each of the five segments of the heating curve for water, posted above. Note changes in potential energy, kinetic energy, phase changes occurring, and changes in intermolecular forces. The first one is done for you.

From -30°C to 0°C (example)

As temperature increases, the average kinetic energy increases, and particles within the solid are now moving faster. There is no phase change, so there is no change in potential energy and no change in intermolecular forces.

a) 0°C to 0°C (melting)

b) 0°C to 100°C (water being heated as a liquid)

c) 100°C to 100°C (vaporizing)

d) beyond 100°C (water being heated as a gas)

For water:

Specific heats (C), J/(g°C):
Ice = 2.11
Water = 4.184
Steam = 2.00

Specific heats (C), J/(moles°C)
Ice = 37.66
Water = 75.38
Steam = 36.57

For water and ice:

Latent Heats (ΔH), J/g:
Heat of fusion: 334
Heat of vaporization: 2260

Latent Heats (ΔH), kJ/mol
Heat of fusion: 6.02
Heat of vaporization: 40.7

3. (2.G, 11.D) How much heat (in J) does it take to vaporize 1.000 L of $H_2O_{(l)}$ already at 100.°C? (Hint: convert the volume to mass, using the density of water, and your SI prefix knowledge.)

4. (2.G, 11.D) How much heat (in J) is released when 10.87 moles of liquid H_2O freezes at 0°C?

5. (2.G, 11.D) How much heat, in kJ, is absorbed while 2.50 moles of water changes from 50.°C to 90.°C?

6. (2.G, 11.D) How much heat, in J, is released while 1.80 moles of steam changes in temperature from 150.°C to 110.°C?

Be careful when temperatures cross melting/freezing/vaporizing/condensation points!

7. (2.G, 11.D) How much heat, in kJ, is released while 25.0 g of water changes from 110.°C to 90.°C?

8. (2.G, 11.D) Calculate the total amount of energy involved to convert 1.00 moles of steam at 100.°C to ice at - 10.°C.

9. (2.G, 11.D) Calculate the total amount of energy involved to convert 10.0 g of solid ice at -20.0°C to 10.0 g of gaseous steam at 140.°C.

Challenge Problems:

10. (3.B, 4.A, 11.A) Substance X (which is *not* water) nevertheless has a heating curve with a similar shape to that of water. It has a melting point of 5°C and a vaporization point of 55°C.

a) Sketch a heating curve for Substance X, using a ruler. Label each axis appropriately, including the unit of measure. Label each segment of the heating curve from A to E starting at the lowest temperature.

b) Answer the following questions related to the heating curve for Substance X.

______In what part(s) of the curve would increasing kinetic energy be displayed?
______In what part(s) of the curve would increasing potential energy be displayed?
______In what part of the curve would the molecules of X be farthest apart?
______In what part of the curve would the molecules of X have the lowest kinetic energy?
______In what part of the curve would the molecules of X have the greatest kinetic energy?

c) Write three relevant, high quality questions about Substance X. Switch course notes with a partner and respond to each other’s questions in writing using the scientific principles explored in this lesson. For question stem ideas, look up “Costa’s House” on Google.

Lesson 6.03 Student Notes:

6.04: Enthalpy of Reaction

1. Use the following thermochemical equation to write new equations that describe each change. ΔH°$_{comb}$ means combustion. ° means "measured under standard conditions".

$C_3H_{8\ (g)} + 5\ O_{2\ (g)} \rightarrow 3\ CO_{2\ (g)} + 4\ H_2O_{\ (g)}$ ΔH°$_{comb}$ - 2220.1 kJ/mol reaction

a) (3.B, 11.C) Reverse the original equation and classify the equation as endothermic or exothermic.

b) (2.G, 3.B, 8.A) Reverse the original equation and rewrite it for the formation of *three* moles of $C_3H_{8\ (g)}$.

2. (2.G, 3.B, 8.E, 8.G) During the process of cellular respiration, your body combusts glucose, $C_6H_{12}O_{6\ (s)}$. The ΔH°$_{comb}$ is - 1273.3 kJ/mol reaction.

Write a balanced thermochemical equation for the combustion of one mole of glucose, including state signifiers and the ΔH°$_{comb}$ statement.

a) Calculate the standard enthalpy when your body burns 100.0 g of glucose.

b) Calculate the standard enthalpy when 500.0 g of water vapor is produced.

c) (Throwback to Unit 5!) What mass of glucose is required to produce 500.0 g of water vapor? (Hint: mass to mass stoichiometry!)

d) The reverse of cellular respiration is (of course) photosynthesis. Calculate the standard enthalpy when a plant produces 200.0 g of glucose.

3. (2.G, 3.B, 8.E, 8.G) Chlorine trifluoride gas reacts with ammonia gas to form nitrogen gas, hydrofluoric acid (in gas form) and chlorine gas. Its ΔH°_{rxn} is - 1196 kJ/mole reaction.

Write a balanced thermochemical equation, including state signifiers and the ΔH°_{rxn} statement.

Calculate ΔH°_{rxn} when 1.00 kg of chlorine trifluoride undergoes this chemical reaction.

4. (2.G, 3.B, 8.E, 8.G) Nitromethane, CH_3NO_2, can be used as a fuel. When the liquid is burned, the (unbalanced) thermochemical equation is mainly:

$CH_3NO_{2\ (l)} + O_{2\ (g)} \rightarrow CO_{2\ (g)} + N_{2\ (g)} + H_2O_{\ (g)}$ $\Delta H°_{rxn} = -1288.5$ kJ/mol reaction

Calculate the enthalpy of reaction to produce 250. g of nitromethane.

5. (2.G, 3.B, 8.E, 8.G) Zinc metal reacts with hydrochloric acid solution in a single replacement reaction, typically in a calorimeter.

Write a balanced equation, including state signifiers

When 0.103 g $Zn_{(s)}$ is combined with enough $HCl_{(aq)}$ to make 50.0 mL of solution, all the zinc reacts, raising the temperature of the solution from 22.5°C to 23.7°C. Determine $\Delta H°_{rxn}$ for this reaction as written. Assume the specific heat and density of the solution is the same as water.

6. (2.G, 3.B, 8.E, 8.G) In a coffee-cup calorimeter, 1.60 g of ammonium nitrate (NH_4NO_3) is mixed with 75.0 g of water at an initial temperature of 25.00℃. After dissolution (dissolving) of the salt, the final temperature of the calorimeter contents is 23.34 ℃.

$NH_4NO_{3\,(s)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$

Assuming the solution has a heat capacity of 4.184 J/℃·g and assuming no heat loss to the calorimeter, calculate the enthalpy change for the dissolution of ammonium nitrate in units of kJ/mole reaction.

Challenge Problems:

7. (2.G, 3.B, 8.E, 8.G) The standard enthalpy of combustion for octane, $C_8H_{18\,(l)}$, is – 5471 kJ/mol reaction.

Write a balanced thermochemical equation for the combustion of octane, including state signifiers and the $\Delta H°_{rxn}$ statement.

a) One gallon of gasoline is 87% octane. Calculate the standard enthalpy of reaction when one (1.0) gallon of gasoline undergoes combustion in a car engine at 25°C. (1 gal = 3.785 L and the density of gasoline is 0.737 g/mL at 25°C. Look up SI prefixes on pg. 11.)

8. (2.G, 3.B, 8.E, 8.G) Consider the dissolution (dissolving) of solid calcium chloride in water:

$CaCl_{2\ (s)} \longrightarrow Ca^{2+}_{\ (aq)} + 2\ Cl^{-}_{\ (aq)}$ $\Delta H^\circ_{rxn} = -\ 81.5$ kJ/mole reaction

An 11.0 g sample of calcium chloride is dissolved in 125 g of water, with both substances initially at 25.0℃. Calculate the final temperature of the solution assuming that the solution has a specific heat capacity of 4.184 J/℃·g. (Hint: Solve for q_p using ΔH°_{rxn}.)

Lesson 6.04 Student Notes:

6.05: Designing a Cold Pack Lab

Introduction

Instant cold packs are used to soothe athletic injuries, pain, and muscle soreness. To operate a cold pack, you break the seal between two compartments, each of which contain chemicals that, when mixed together, undergo an endothermic change. The overall enthalpy of this change would be a positive value as energy is absorbed by the system (the chemicals) and lost by the surroundings.

The simplest cold pack involves dissolving a salt (an ionic compound) in water, forming a solution. To measure the **enthalpy of solution**, chemists use a calorimeter. On a high school budget, the calorimeter we would use is a nested pair of foam coffee cups, known as a coffee cup calorimeter. A predetermined volume of water is put into one of the cups and the initial and final temperature of the water is measured as a measured mass of salt is added to the water. Mixing the salt in the water is essential for thorough dissolution (dissolving). We would then use the formula $q = mC\Delta T$ to determine the energy change in the water.

The First Law of Thermodynamics states that the total energy of the universe is constant, so the energy change in the water (the surroundings) is equal to the energy change of the dissolving salt (the system). Since enthalpy depends on the amount of salt used, we would then calculate the **molar enthalpy of solution,** or $\boldsymbol{\Delta H_{soln}}$ (units kJ/mol) by dividing the enthalpy we determined by the number of moles of the salt that was used.

Your Task

You will act as a small research group contracted by another company who seeks to create an economical, yet effective, cold pack. The cold pack will be small and portable, containing two compartments, one filled with water and the other containing an ionic compound. The water-filled compartment will be inside the salt-filled compartment so that when the compartments are broken the salt and water will mix together. 60 mL of water will be used in each cold pack and the temperature must fall by about 2 ℃ to be effective. (Not less, and not more.)

To determine which is the most economical, you will also have to determine how much of each type of salt you will need to use and how much it will cost per cold pack. You will investigate three different ionic compounds for potential use in this cold pack system. Using the data you collect, along with cost data (provided below), determine which of the three salts should be used to produce an effective but economical cold pack when a 3-4 g sample of salt dissolves in water.

Effectiveness is measured in terms of **enthalpy of solution**, measured in units of kJ/mol reaction (of salt). The more heat lost during the process of dissolving (or **dissolution**), the colder the temperature of the water will be. Follow all lab safety rules and read the SDS for each chemical. Once you have gathered all data in multiple trials, you will then have to show your work and calculate ΔH_{soln} for each salt, then compare the values to determine the most effective salt to use in a cold pack. The salts, and their costs per unit, are in the table below.

Salt (name and chemical formula)	**Cost / Unit**
$NH_4Cl_{(s)}$ (ammonium chloride)	$13.90 / 1000 g
$MgSO_4 \cdot 7\ H_2O$ (s) (magnesium sulfate heptahydrate)	$1.17 / 100 g
$Na_2S_2O_3 \cdot 5\ H_2O$ (s) (sodium thiosulfate pentahydrate)	$8.55 / 500 g

In the end, you can only choose one of the three salts. Write a well-written group conclusion. Your argument must include a claim, which is your answer to the guiding question, and include evidence (your data for all trials, all your calculations, and your data analysis) in support of your claim. Lastly, you must include a reasoning section where you justify the evidence in your argument. You will need to use a scientific concept or principle to explain why the evidence you decided to use is relevant and important. Choose well!

Adapted from: "Lab 28. Designing a Cold Pack: Which Salt Should Be Used to Make an Effective but Economical Cold Pack?" *Argument-Driven Inquiry in Chemistry: Lab Investigations for Grades 9-12*, by Victor Sampson, NSTA Press, National Science Teachers Association, 2015, pp. 448–452.

Lesson 6.05 Student Notes:

6.06: Hess' Law

1. (3.B) Describe a "real life scenario" (not related to travel routes!) that could be a state function.

2. (2.G, 3.B, 11.C) Calculate the enthalpy of reaction for: $PbCl_{2\,(s)} + Cl_{2\,(g)} \rightarrow PbCl_{4\,(l)}$

by combining: $Pb_{(s)} + Cl_{2\,(g)} \rightarrow PbCl_{2\,(s)}$ $\Delta H = -359.4\ kJ$

$Pb_{(s)} + 2\,Cl_{2\,(g)} \rightarrow PbCl_{4\,(l)}$ $\Delta H = -329.3\ kJ$

3. (2.G, 3.B, 11.C) From the following heats of reaction:

$2\,SO_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,SO_{3\,(g)}$ $\Delta H = -196\ kJ$
$2\,S_{(s)} + 3\,O_{2\,(g)} \rightarrow 2\,SO_{3\,(g)}$ $\Delta H = -790\ kJ$

Calculate the heat of reaction for: $S_{(s)} + O_{2\,(g)} \rightarrow SO_{2\,(g)}$

4. (2.G, 3.B, 11.C) From the following heats of reaction:

$2\ H_{2\ (g)} + O_{2\ (g)} \rightarrow 2\ H_2O_{\ (g)}$ ΔH = - 483.6 kJ
$3\ O_{2\ (g)} \rightarrow 2\ O_{3\ (g)}$ ΔH = 284.6 kJ

Calculate the enthalpy of reaction for: $3\ H_{2\ (g)} + O_{3\ (g)} \rightarrow 3\ H_2O_{\ (g)}$

5. (2.G, 3.B, 11.C) Calculate the enthalpy for: $N_{2\ (g)} + 2\ O_{2\ (g)} \rightarrow 2\ NO_{2\ (g)}$

by combining:
$N_{2\ (g)} + O_{2\ (g)} \rightarrow 2\ NO_{\ (g)}$ ΔH° = + 180 kJ
$2\ NO_{2\ (g)} \rightarrow 2\ NO_{\ (g)} + O_{2\ (g)}$ ΔH° = + 112 kJ

6. (2.G, 3.B, 11.C) Find the ΔH for the reaction below, given the following intermediate reactions and subsequent ΔH values:

$N_2H_{4\,(l)} + H_{2\,(g)} \rightarrow 2\ NH_{3\,(g)}$ ΔH = ???

$N_2H_{4\,(l)} + CH_4O_{\,(l)} \rightarrow CH_2O_{\,(g)} + N_{2\,(g)} + 3\ H_{2\,(g)}$ ΔH = - 37 kJ

$N_{2\,(g)} + 3\ H_{2\,(g)} \rightarrow 2\ NH_{3\,(g)}$ ΔH = - 46 kJ

$CH_4O_{\,(l)} \rightarrow CH_2O_{\,(g)} + H_{2\,(g)}$ ΔH = - 65 kJ

7. (2.G, 3.B, 11.C) Find the ΔH for the reaction below, given the following intermediate reactions and subsequent ΔH values:

$\frac{1}{2} H_{2\,(g)} + \frac{1}{2} Cl_{2\,(g)} \rightarrow HCl_{\,(g)}$ ΔH = ???

$COCl_{2\,(g)} + H_2O_{\,(l)} \rightarrow CH_2Cl_{2\,(l)} + O_{2\,(g)}$ ΔH = 47.5 kJ

$2\ HCl_{\,(g)} + \frac{1}{2} O_{2\,(g)} \rightarrow H_2O_{\,(l)} + Cl_{2\,(g)}$ ΔH = 105 kJ

$CH_2Cl_{2\,(l)} + H_{2\,(g)} + \frac{3}{2} O_{2\,(g)} \rightarrow COCl_{2\,(g)} + 2\ H_2O_{\,(l)}$ ΔH = - 402.5 kJ

Challenge Problems:

8. (2.G, 3.B, 11.C) One reaction involved in the conversion of iron ore to the metal is:

$$FeO_{(s)} + CO_{(g)} \rightarrow Fe_{(s)} + CO_{2\,(g)}$$

Calculate the standard enthalpy change for this reaction from the following intermediate reactions of iron oxides with carbon monoxide:

$3\ Fe_2O_{3\,(s)} + CO_{(g)} \rightarrow 2\ Fe_3O_{4\,(s)} + CO_{2\,(g)}$ $\Delta H = -47$ kJ

$Fe_2O_{3\,(s)} + 3\ CO_{(g)} \rightarrow 2\ Fe_{(s)} + 3\ CO_{2\,(g)}$ $\Delta H = -25$ kJ

$Fe_3O_{4\,(s)} + CO_{(g)} \rightarrow 3\ FeO_{(s)} + CO_{2\,(g)}$ $\Delta H = 19$ kJ

9. (2.G, 3.B, 11.C) Find the ΔH for the reaction below, given the following intermediate reactions and subsequent ΔH values:

$$C_2H_{2\,(g)} + \frac{5}{2} O_{2\,(g)} \rightarrow 2\ CO_{2\,(g)} + H_2O_{(g)} \quad \Delta H = ???$$

$C_2H_{6\,(g)} \rightarrow C_2H_{2\,(g)} + 2\ H_{2\,(g)}$ $\Delta H = 283.5$ kJ

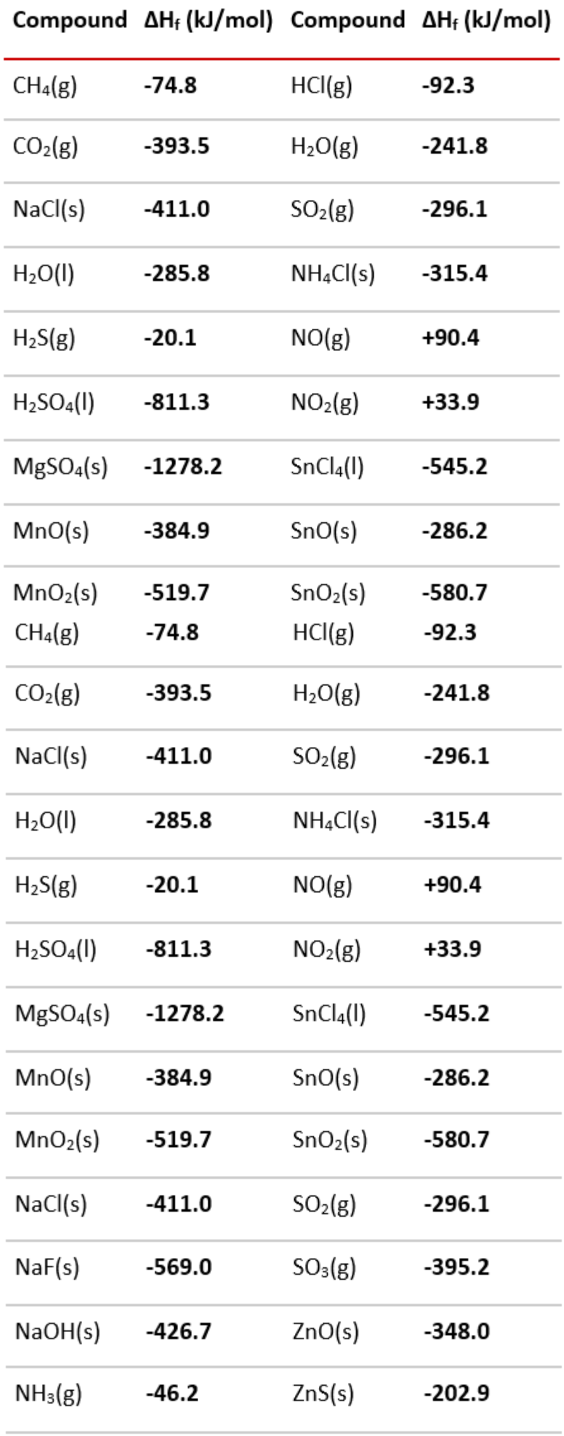
$H_{2\,(g)} + \frac{1}{2} O_{2\,(g)} \rightarrow H_2O_{(g)}$ $\Delta H = -213.7$ kJ

$2\ CO_{2\,(g)} + 3\ H_2O_{(g)} \rightarrow C_2H_{6\,(g)} + \frac{7}{2} O_{2\,(g)}$ $\Delta H = +849$ kJ

Lesson 6.06 Student Notes:

6.07: Enthalpy of Formation and Energy Diagrams

(2.G, 3.B, 11.C) Use the standard enthalpies of formation table to determine the change in enthalpy for each of these reactions.



Compound	ΔH_f (kJ/mol)	Compound	ΔH_f (kJ/mol)
CH_4(g)	**-74.8**	HCl(g)	**-92.3**
CO_2(g)	**-393.5**	H_2O(g)	**-241.8**
NaCl(s)	**-411.0**	SO_2(g)	**-296.1**
H_2O(l)	**-285.8**	NH_4Cl(s)	**-315.4**
H_2S(g)	**-20.1**	NO(g)	**+90.4**
H_2SO_4(l)	**-811.3**	NO_2(g)	**+33.9**
$MgSO_4$(s)	**-1278.2**	$SnCl_4$(l)	**-545.2**
MnO(s)	**-384.9**	SnO(s)	**-286.2**
MnO_2(s)	**-519.7**	SnO_2(s)	**-580.7**
CH_4(g)	**-74.8**	HCl(g)	**-92.3**
CO_2(g)	**-393.5**	H_2O(g)	**-241.8**
NaCl(s)	**-411.0**	SO_2(g)	**-296.1**
H_2O(l)	**-285.8**	NH_4Cl(s)	**-315.4**
H_2S(g)	**-20.1**	NO(g)	**+90.4**
H_2SO_4(l)	**-811.3**	NO_2(g)	**+33.9**
$MgSO_4$(s)	**-1278.2**	$SnCl_4$(l)	**-545.2**
MnO(s)	**-384.9**	SnO(s)	**-286.2**
MnO_2(s)	**-519.7**	SnO_2(s)	**-580.7**
NaCl(s)	**-411.0**	SO_2(g)	**-296.1**
NaF(s)	**-569.0**	SO_3(g)	**-395.2**
NaOH(s)	**-426.7**	ZnO(s)	**-348.0**
NH_3(g)	**-46.2**	ZnS(s)	**-202.9**

1. $NaOH_{(s)} + HCl_{(g)} \rightarrow NaCl_{(s)} + H_2O_{(g)}$

2. $2\ CO_{(g)} + O_{2\,(g)} \rightarrow 2\ CO_{2\,(g)}$ ($\Delta H°f$ for $CO_{(g)}$ = -110.5 kJ/mol)

3. $CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(l)}$

4. $2\ H_2S_{(g)} + 3\ O_{2\,(g)} \rightarrow 2\ H_2O_{(l)} + 2\ SO_{2\,(g)}$

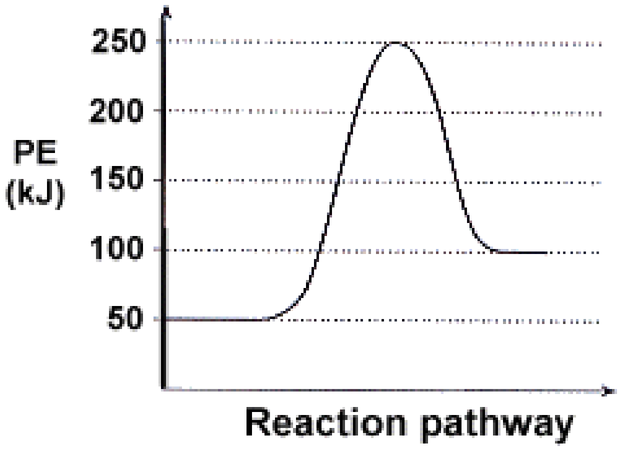
5. $H_2SO_{4\,(l)} + 2\ NaOH_{(s)} \rightarrow Na_2SO_{4\,(aq)} + 2\ H_2O_{(l)}$ ($\Delta H°f$ for $Na_2SO_{4\,(aq)}$ = -1387 kJ/mol)

6. (2.E, 3.B) Marble chips are pure calcium carbonate, $CaCO_{3\,(s).}$ Calcium carbonate reacts quite well with hydrochloric acid, $HCl_{\,(aq)}$. Design a procedure that will cause the rate of reaction for 10. g of calcium carbonate with hydrochloric acid to be as fast as possible.

You have available: marble chips, hydrochloric acid of different concentrations (low and high), a thermometer, a hot plate, beakers, ice, tap water, a mortar and pestle, tongs, and a stopwatch. (Hint: Not all these items have to be used!)

7. (2.G, 3.B, 11.C) Given the chemical equation for a reaction, A + B → C (all questions are for the forward reaction):

[12]



a) Does the energy diagram illustrate an exothermic or an endothermic reaction? State one reason, in terms of energy, to support your answer.

b) Draw a labelled line on the energy diagram that represents the activation energy of the reaction. What is the value of the activation energy in kilojoules?

c) Draw a labelled line on the energy diagram for ΔH_{rxn}. What is the value of ΔH_{rxn}?

d) Draw a dashed line on the diagram to indicate an energy diagram for the reaction if a catalyst is added.

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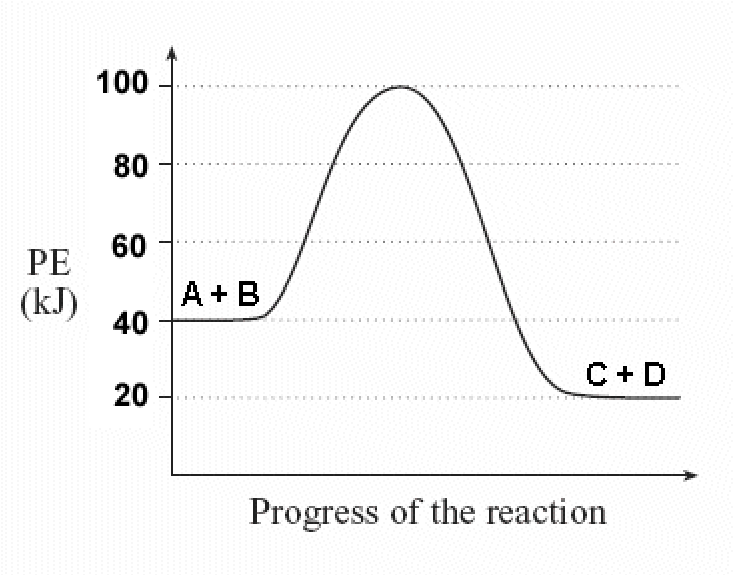
[12] Lawson, Peggy. “Chemical Kinetics.” *Chemistry 30*, Saskatchewan Learning, 2004, sites.prairiesouth.ca/legacy/chemistry/chem30/2_kinetics/practice/q3_2.htm. This site is copyright free and available for anyone's use.

8. In the space below, sketch a labelled energy diagram for the reverse reaction, C ⟶ A + B. Use the information above to help you.

(2.G, 3.B, 11.C) Answer all questions related to this energy diagram below:

[13]



a) Is the *reverse* reaction endothermic or exothermic? State one reason, in terms of energy, to support your answer.

b) Is the activation energy, **E_a**, for the forward reaction, less than, equal to, or greater than to the activation energy for the reverse reaction? Support your answer with data and calculations.

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[13] Lawson, Peggy. "Chemical Kinetics." *Chemistry 30*, Saskatchewan Learning, 2004, sites.prairiesouth.ca/legacy/chemistry/chem30/2_kinetics/practice/q3_2.htm. This site is copy-right free and available for anyone's use.

c) Is the enthalpy of reaction, $\Delta H°_{\text{rxn}}$, for the forward reaction, less than, equal to, or greater than the enthalpy of reaction for the reverse reaction? Support your answer with data and calculations.

Challenge Questions:

9. (2.G, 3.B, 11.C) Sketch a labelled energy diagram represented by the following values of ΔH and E_a. You may make up appropriate values for the y-axis. Be sure to label both axes.

a) ΔH_{rxn} = - 100 kJ and E_a = 20 kJ (forward), for the reaction A + B C

b) ΔH_{rxn} = + 60 kJ and E_a = 40 kJ (forward), for the reaction D + E F

10. Which value, $\Delta H°_{\text{rxn}}$ or E_a, would be more useful in predicting the kinetics (speed) of a reaction? Explain your answer.

11. (3.B) Describe a situation where it would be useful to slow down the rate of a reaction. This could be from your own experience or from a scientist's perspective. In your description, describe what could be done to slow down the rate and why this would be useful.

Lesson 6.07 Student Notes:

Unit 7: Gases

7.01: Kinetic Molecular Theory (KMT) and Gas Law Variables

1. (2.G) The pressure in a room is recorded as 738 mmHg. Convert this measurement to atm.

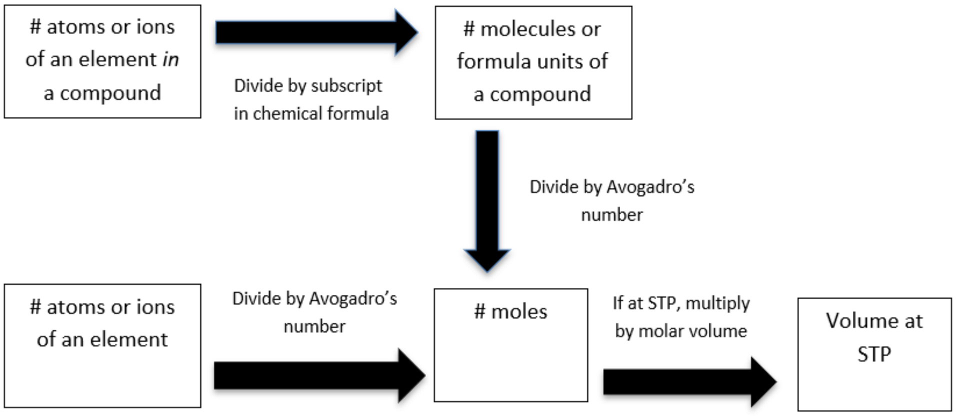
2. (2.G) A ball is inflated to a pressure of 32.0 pounds per square inch (psi). Convert this pressure to kilopascals (kPa). (14.7 psi is equal to 1 atm.)

3. (2.G) Convert the following temperature values.

10.0°C = ___________ K
1111 K = ___________°C
25°C = ___________ K
283 °C = ___________ K
323 K = ___________ °C
367 K = ___________ °C

The Mole Map for Gases

Find out the quantity the problem is giving you to start with and follow the arrows. If you have to go in the reverse direction of the arrows, do the opposite operation. Convert between moles and mass by using molar mass as a conversion factor.


# atoms or ions of an element *in* a compound
Divide by subscript in chemical formula
# molecules or formula units of a compound
Divide by Avogadro's number
# atoms or ions of an element
Divide by Avogadro's number
# moles
If at STP, multiply by molar volume
Volume at STP

4. (2.G) Calculate the volume, in liters, occupied by 0.030 moles of a gas at STP.

5. (2.G, 8.B) How many oxygen molecules are in 3.36 L of oxygen gas at STP?

6. (2.G, 8.B) How many nitrogen atoms are in 35.0 L of nitrogen dioxide gas at STP?

Challenge Questions

7. (2.G, 9.B) Which has a larger volume at STP, 3.00×10^{24} $N_{2\,(g)}$ molecules or 4.00×10^{23} $CO_{2(g)}$ molecules? Do this two ways: first, by using a kinetic molecular theory based explanation, then, by using calculations.

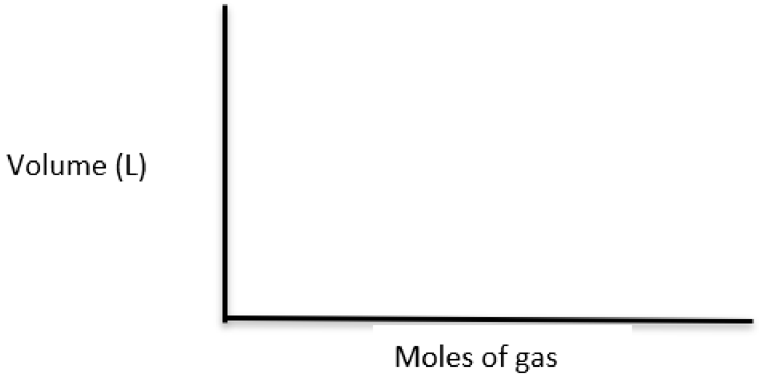
8. Which has a larger *mass* at STP, 6.022×10^{23} $N_{2\,(g)}$ molecules or 6.022×10^{23} $CO_{2\,(g)}$ molecules? Use the concept of molar mass to explain your answer.

9. (2.G, 2.H, 9.B) Using scientific principles such as force, area, and the kinetic molecular theory, explain why high-altitude situations (like at the tops of mountains) are always low pressure environments.

Lesson 7.01 Student Notes:

7.02: Gas Laws: Avogadro, Boyle, Charles, Gay-Lussac, and Combined Gas Law

1. (2.H, 9.A) On the axes that follow, draw a line representing the directly proportional relationship between volume and moles of gas at constant temperature and pressure.

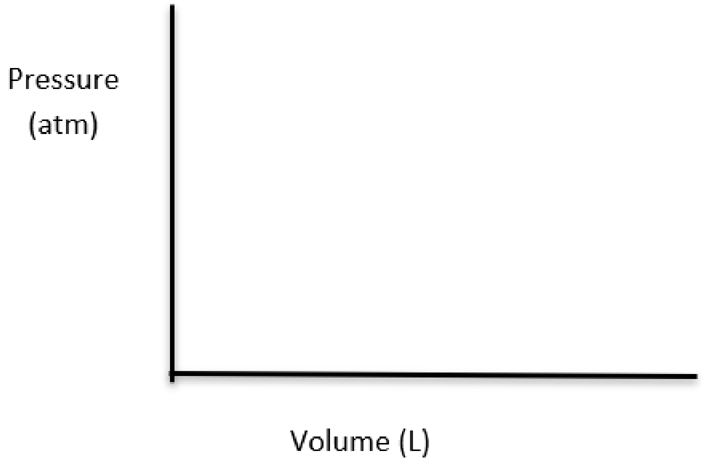


2. (2.G, 8.A, 9.A) 5.00 L of a certain gas is known to contain 0.965 moles. If 0.835 moles of gas are added to that sample, what will be the new volume, assuming an unchanged temperature and pressure?

3. (2.G, 8.A, 9.A) 13.1 g of $O_{2\,(g)}$ is in a 4.37 L container at constant pressure. The volume of the container decreases to 2.72 L. How many grams of gas remain? (Hint: first, how will you convert the given $O_{2\,(g)}$ mass to moles?)

4. (2.G, 9.A) Answer this question without doing calculations: If the number of moles of gas doubles inside a container (assuming Avogadro's law holds), what would be the corresponding change in volume? Explain your answer.

5. (2.H, 9.A) On the axes that follow, draw a line representing the indirectly proportional relationship between pressure and volume at constant temperature and number of moles.



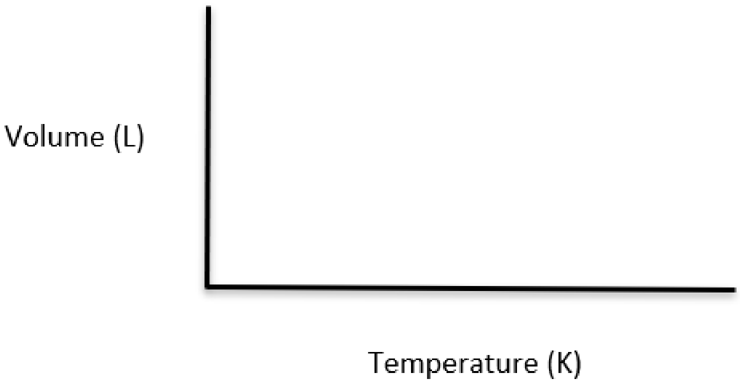
(FYI: the actual graph for this relationship does have a curve in it, but the trend is still the same.)

6. (2.G, 9.A) Consider a 1.53 L sample of a gaseous SO_2 at a pressure of 5.6×10^3 Pa. If the pressure is changed to 1.5×10^4 Pa at constant temperature, what will be the new volume of gas?

7. (2.G, 9.A) A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L. If the balloon is filled with 2.0 L of helium at sea level (1 atm), is released, and rises to an altitude at which the atmospheric pressure is only 500. mmHg, will the balloon burst? Assume temperature is constant.

8. (2.H, 9.A) Answer this question without doing calculations: If the pressure of the gas inside a balloon triples, what will happen to the corresponding volume, assuming Boyle's law holds? Explain your answer.

9. (2.H, 9.A) On the axes that follow, draw a line representing the directly proportional relationship between volume and temperature at constant pressure and number of moles.

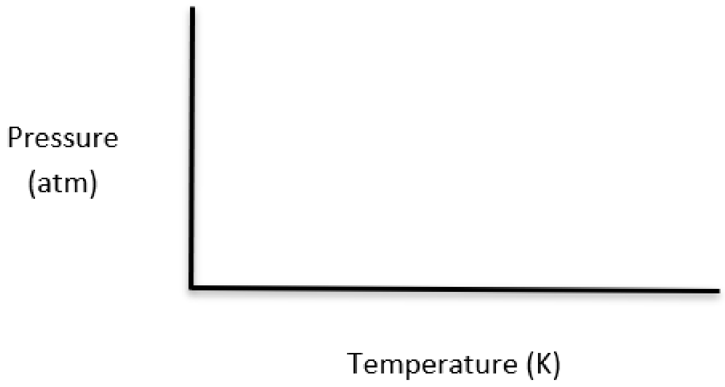


10. (2.G, 9.A) A gas with a volume of 600. mL has a temperature of 303 K. At constant pressure the gas is heated until the gas doubles in volume. What is the new temperature of the gas in Celsius?

11. (2.G, 9.A) On hot days, you may have noticed that potato chip bags seem to “inflate”, even though they have not been opened. If I have a 250. mL bag at a temperature of 30.0°C, and I leave it in my car which has a temperature of 40.0°C on a very hot day, what will the new volume of the bag be?

12. (2.G, 9.A) Answer this question without doing calculations: If the volume of the gas inside a balloon decreases by 2 (ie. it is now half of what it was), what will happen to the corresponding temperature, assuming Charles’s law holds? Explain your answer.

13. (2.H, 9.A) On the axes that follow, draw a line representing the directly proportional relationship between pressure and temperature at constant temperature and number of moles.



14. (2.G, 9.A) A gas sample is found to exert 97 kPa of pressure at 25.0°C. What would be the required temperature (in Celsius) to change the pressure to standard pressure?

15. (2.G, 9.A) A 28.4 L sample of nitrogen gas inside a rigid, metal container at 51.0°C is placed inside an oven whose temperature is 254°C. The pressure inside the container at 51.0°C was 2.7 atm. What is the pressure of the nitrogen gas in kPa after the temperature is increased?

16. (2.G, 9.A) Answer this question without doing calculations: If the temperature of the gas inside a balloon decreases by 2 (i.e. it is now half of what it was), what will happen to the corresponding pressure, assuming Gay-Lussac's law holds? Explain your answer.

17. (2.G, 9.A) A gas is heated from 263.0 K to 298.0 K and the volume is increased from 24.0 liters to 35.0 liters by moving a large piston within a sealed cylinder. If the original pressure was 1.00 atm, what would the final pressure be?

18. (2.G, 9.A) The pressure of a gas is reduced from 1200.0 mm Hg to 850.0 mm Hg as the volume of its container is increased by moving a piston from 85.0 mL to 350.0 mL. What would the final temperature be if the original temperature was 90.0 °C?

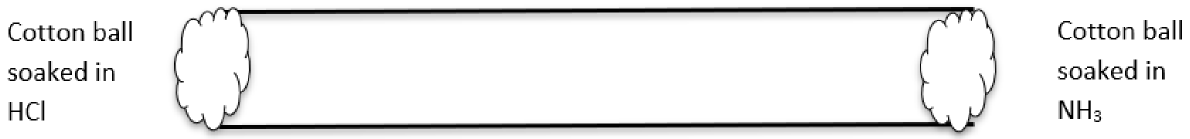
Challenge Problem:

19. (2.G, 9.A) Write an original Combined Gas Law problem of your own design, solving for one of the four gas variables if the other three values change. You may either solve it yourself or solve someone else's problem. Don't forget to show all work and use sig fig rules in your final answer.

Lesson 7.02 Student Notes:

7.03: Graham's Law of Diffusion and Effusion

1. (3.B, 9.B) The classic example of the relationship between molar mass and diffusion/effusion involves a long glass tube with soaked cotton at either end. One cotton ball contains hydrochloric acid (HCl) and the other contains ammonia (NH_3). Each liquid emits a vapor.



The HCl $_{(g)}$ and $NH_{3\,(g)}$ will diffuse from their respective cotton balls towards each other in the main area of the tube and form a white solid ring of ammonium chloride, NH_4Cl (s).

Where would the ring be formed: closer to the HCl, closer to the NH_3, or right in the middle between the two cotton balls? Explain your answer in terms of the kinetic molecular theory, and diffusion concepts.

2. (3.B, 3.F, 9.B) Explain how Graham's law relates to the postulates of the kinetic molecular theory.

3. (2.G, 3.B) Under the same conditions of temperature and pressure, how many times faster will hydrogen gas effuse, compared to carbon monoxide?

4. (2.G, 3.B) If the carbon monoxide in the previous problem takes 32 seconds to effuse, how long will the hydrogen take to effuse?

5. (2.G, 3.B) At a certain temperature, O_2 molecules move with an average velocity of 345 mph. At that same temperature, what would be the average velocity of:

a) He atoms?

b) CO_2 molecules?

6. (2.G, 3.B) At a certain temperature, CH_4 molecules move with an average velocity of 187 m/sec. At that same temperature, gas X particles have an average velocity of 141 m/sec.

a) Is gas X heavier or lighter than CH_4?

b) What is the molecular weight (molar mass) of gas X?

Lesson 7.03 Student Notes:

7.04: Ideal Gas Law; Density and Molar Mass of a Gas

1. (3.B, 9.A) Let's say that a sample of hydrogen gas at 600. K and 0.01 atm is acting as an ideal gas. Would each given sample of gas be more ideal, or more real, than the hydrogen sample? Explain each of your answers.

a) $CO_{2\ (g)}$ at 600. K and 0.01 atm

b) $H_{2\ (g)}$ at 100. K and 250 atm

2. (3.B, 9.A) Since R = R, show that you can derive the Combined Gas Law using the Ideal Gas Law formula.

3. (2.G, 8.A, 9.A) Calculate the temperature in °C of 0.50 g of nitrogen gas occupying a 15.0 L rigid tank at a pressure of 650. mmHg.

4. (2.G, 8.B, 9.A) How many moles of helium gas are trapped in a 25.0 L balloon when the gas has a pressure of 102 kPa and a temperature of 9.00℃?

a) How many molecules is this?

5. (2.G, 8.A, 9.A). What is the mass of oxygen gas in a 80.0 L balloon at room temperature, 20.0℃, and a pressure of 99.3 kPa?

6. (2.G, 8.A, 9.A). Calculate the pressure in atm of 10.0 g of propane gas in a cylinder with a volume of 50.0 L and a temperature of 35.0℃. (Propane is C_3H_8.)

7. (2.G, 8.A, 9.A) An oxygen gas sample has 2.50 g of gas at a pressure of 0.900 atm and is held at 19.0℃. What is the volume of the sample?

8. (2.G, 8.A, 9.A) What is the volume of 1.00 g of butane gas at STP? (Butane is C_4H_{10}.)

9. (2.G, 8.A, 9.A) Determine the density of CO_2 at 745 mmHg and 65.0°C.

10. (2.G, 8.A, 9.A) An unknown gas has a density of 3.167 g/L at STP. What is the identity of the gas? (Possible choices: Ar, O_2, Cl_2, HF, or H_2O?)

11. (2.G, 8.A, 9.A) Calculate the density a gas will have at STP if its molar mass is 28.7 g/mol.

12. (2.G, 8.A, 9.A) Uranium hexafluoride, UF_6, is a solid at room temperature (25.0^o C), but it boils at 56.0^{o}C. Determine the density of uranium hexafluoride at 60.0^{o}C and 745 mmHg.

13. (2.G, 8.A, 9.A) Calculate the molar mass of a gas if 35.44 g of the gas stored in a 7.50 L tank exerts a pressure of 60.0 atm at a constant temperature of 35.5 °C.

Challenge Problems:

14. (2.G, 8.A, 8.D, 9.A) A gaseous hydrogen and carbon containing compound is decomposed and found to contain 82.66% carbon and 17.34% hydrogen by mass. The mass of 158 mL of the gas, measured at 556 mmHg and 25.0°C, is 0.275 g. What is the molecular formula of the compound?

Lesson 7.04 Student Notes:

7.05: Dalton's Law of Partial Pressures / Ideal Gas Law Lab

1. (2.G, 9.A) A tank containing ammonia and argon has a total pressure equal to 1.8 atm. The pressure of the ammonia is 1.2 atm. What is the pressure of the argon gas?

2. (2.G, 8.A) Calculate the mole fraction of each gas in a tank containing 5.00 g each of oxygen, carbon dioxide, and hydrogen gas.

a) Sketch a particle diagram of the tank. Use different symbols for each gas. You do not have to show the relative amounts of each gas "to scale", but it should be clear which gas or gases make up the majority of the tank.

b) (2.G, 9.A) Calculate the partial pressure of each gas, if the total pressure of the gas mixture is 110. kPa.

3. (2.G, 9.A) The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of nitrogen in air when the atmospheric pressure is 760. torr.

4. (2.G, 8.A, 9.A) A 1.00-L mixture of helium, neon, and argon has a total pressure of 662 mmHg at 298 K. If the partial pressure of helium is 341 mmHg and the partial pressure of neon is 112 mmHg, what is the *mass* of argon present in the mixture?

5. (2.G, 8.A, 9.A) A 12.5-L scuba diving tank contains a helium-oxygen (heliox) mixture of 24.2 g He and 4.32 g of O_2 at 25.0°C. Calculate the mole fraction and partial pressure of each component in the mixture and the total pressure of the tank.

6. (2.G, 8.A, 9.A) For a particular dive, 46 L of helium and 12 L of oxygen, both at 25.0°C and 1 atm, are pumped into a tank with a volume of 5.0 L. Calculate the partial pressure of each gas and the total pressure *in the tank* under these conditions.

7. (2.G, 8.A, 9.A) 193 mL of O_2 was collected over water on a day when the atmospheric pressure was 762 mmHg. The temperature of the water was 23.0°C and at this temperature the vapor pressure of the water is 21.1 torr.

a) How many grams of oxygen were collected?

b) How many moles of water are in the gas mixture? (Hint: mole fraction)

Be careful - sometimes you can exert changes on the gas you collect!

8. (2.G, 8.A, 9.A) An experiment generates 2.58 L of hydrogen gas collected over water. The temperature is 20. °C and the atmospheric pressure is 98.6 kPa. Calculate the volume of the dry hydrogen gas *at STP* if the vapor pressure of water at 20.0 °C is 17.54 mmHg.

9. (2.H, 8.A, 9.A) A gas is collected over water and occupies a volume of 596 mL at 43.0°C. The atmospheric pressure is 101.1 kPa. What volume will the dry gas occupy at this temperature and at standard pressure? (The vapor pressure of water at this temperature is 8.6 kPa.)

Ideal Gas Law Lab

Problem: In this lab activity you will collect a sample of carbon dioxide gas over water and experimentally determine the molar mass and density of carbon dioxide. We will dissolve part of an antacid tablet (Alka Seltzer(c) type) to generate the gas.

Procedure

Pneumatic trough or large plastic tub
Glass tubing
250 ml Erlenmeyer flask
Stopper
Rubber tubing
100 mL graduated cylinder
Antacid tablet
Electronic balance

Partner 1 (or half of your lab group)

1. Fill your pneumatic trough with tap water, making sure the water level rises above the metal tray set in the trough.

2. Fill your graduated cylinder all the way to the top with tap water. Put your hand over the graduated cylinder. Holding your hand firmly over the top, invert the graduated cylinder and submerge it in the trough. You will get wet!

3. Once submerged, remove your hand, and slide the graduated cylinder onto the tray. You may need to lift the graduated cylinder a bit to do so. Your graduated cylinder should still be completely full of water. You will probably have to hold the graduated cylinder to ensure it stays still.

4. Make sure your glass bend tightly fits into the rubber tubing so no gas escapes. Insert the glass tubing halfway into the graduated cylinder through the hole in the metal tray. Make sure the other end of the rubber tubing fits tightly onto the arm of the Erlenmeyer flask.

Partner 2 (or the other half of your lab group)

5. Add tap water to a 250 mL Erlenmeyer flask with arm until it is approximately half full. You will need to experiment to determine the amount of water to antacid weight that is needed to generate 80-90 mL of gas.

6. Find the mass of the flask and tap water. Record your data in the table below.

7. Find the mass of of an antacid tablet. You will need to experiment using one of your trials to determine the amount of antacid tablet needed in order to make about 80-90 mL of collected gas. Record your data in the table below.

Whole Group (after steps 1-7 are complete)

8. Add the antacid tablet to the flask and **quickly cap the flask with the rubber stopper.**

9. Once the flask stops fizzing, and the liquid level in the bottle stops moving, remove the glass bend from the cylinder. BE CAREFUL NOT TO TIP IT OVER!!

10. Slide the cylinder off the tray and slowly raise or lower it so the liquid level inside matches the level in the trough. This is to ensure that the total pressure of gas inside the cylinder is equal to the atmospheric pressure outside of it. Be careful not to pull it out of the trough completely!

11. Once the water levels match remove the cylinder from the trough.

12. Find the mass of the flask and its remaining contents. Record the mass.

13. Record the room temperature and atmospheric pressure.

Data

Volume Data:

Measurement	**Trial 1**	**Trial 2**	**Trial 3**
Volume of carbon dioxide gas (mL)			

Mass Data:

Measurement	**Trial 1**	**Trial 2**	**Trial 3**
Mass of flask + water (g)			
Mass of antacid tablet (g)			
Total mass of reactants (g)			
Mass of flask + contents remaining (g)			
Mass of carbon dioxide gas (g)			

Pressure and Temperature Data:

To obtain atmospheric pressure, you may either consult the local weather for “barometric pressure” or download and use the Google Science Journal app for your phone. Either way, you will need to convert that measurement. Use online sources to look up the water vapor pressure of H_2O at room temperature.

Atmospheric Pressure (mmHg)	
Partial Pressure of H_2O (mmHg) at room temperature:	
Water Temperature (°C)	

Calculations:

a) Calculate the molar mass of CO_2 from your lab data in the space below for all trials. Show all work!

b) Calculate your average percent error for the molar mass of carbon dioxide.

c) Calculate the density of carbon dioxide at these conditions.

Analysis: Describe the effect of each source of error, or mistake, on your apparent (calculated) value for the molar mass of carbon dioxide - does it make the molar mass too large or too small? Explain how the error, or mistake, affects one of the variables and how that affects your apparent molar mass.

a) The tube slipped out from under the graduated cylinder while the carbon dioxide gas was produced.

b) You used the atmospheric pressure in your calculation for molar mass. (i.e. you did not subtract the vapor pressure of water first, like you were supposed to.)

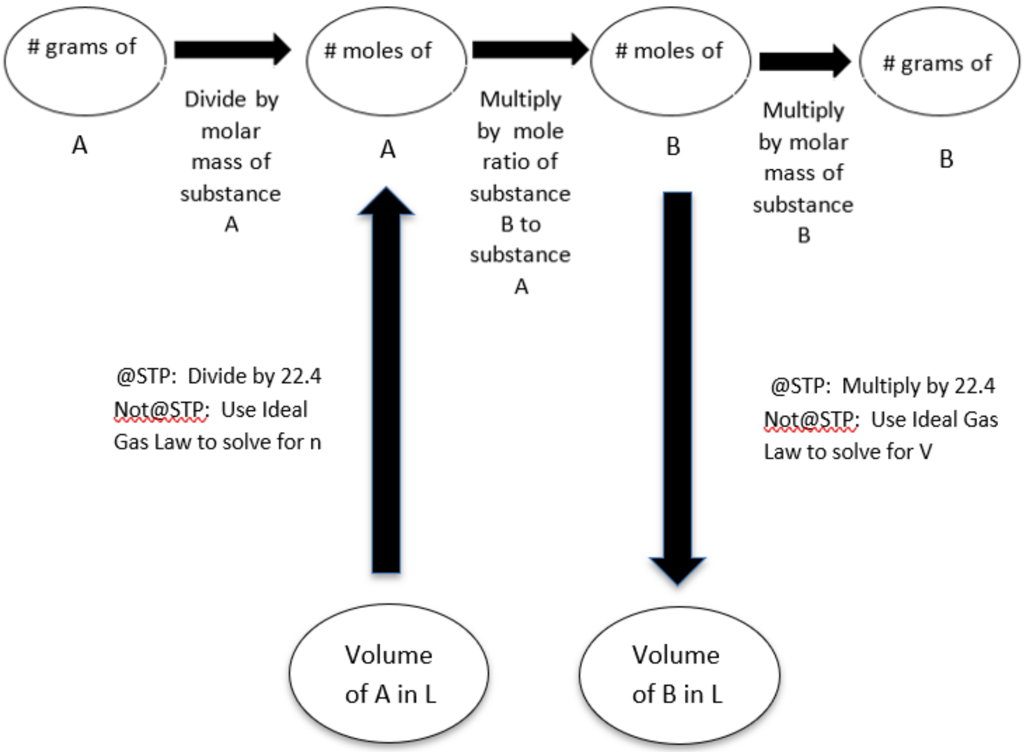
c) You forgot to convert the temperature from Celsius to Kelvin.

7.06: Gas Stoichiometry

1. (2.G, 8.A, 9.A) What volume would be occupied by 2.0 mol nitrogen gas at 0°C and 1 atm?

2. (2.G, 8.A, 9.A) What volume would be occupied by 88.0 g of gaseous carbon monoxide at STP?

Use this to convert from substance A (given quantity) to substance B (desired quantity). You MUST have a balanced chemical equation first.



# grams of
A
Divide by molar mass of substance A
# moles of
A
Multiply by mole ratio of substance B to substance A
# moles of
B
Multiply by molar mass of substance B
# grams of
B
@STP: Divide by 22.4
Not@STP: Use Ideal Gas Law to solve for n
@STP: Multiply by 22.4
Not@STP: Use Ideal Gas Law to solve for V
Volume of A in L
Volume of B in L

3. Propane tanks are used for cooking and heating. When propane (C_3H_8) burns, the products of the reaction are carbon dioxide and water vapor.

Write the balanced chemical equation here, including the state signifiers.

a) At STP, how many liters of oxygen would be needed to completely combust 0.350 L of propane? (Hint: Start by writing a balanced chemical equation for the combustion of propane.)

b) (2.G, 8.A, 8.G) How many liters of water vapor would be produced by the reaction of 23.0 g of propane at STP?

c) (2.G, 8.A, 8.G, 9.A) Calculate the theoretical yield, in L, of carbon dioxide if 1.00 kg of propane burns with excess oxygen at a pressure of 102 kPa and temperature of 300.°C.

4. A sample of solid potassium chlorate was heated in a test tube and decomposes into solid potassium chloride and oxygen gas.

Write the balanced chemical equation here, including the state signifiers.

The oxygen produced was collected over water at 22℃ at an atmospheric pressure of 754 torr. The volume of the gas collected was 0.650 L, and the vapor pressure of water at this temperature is 21 torr.

a) (2.G, 9.A) Calculate the partial pressure of oxygen in the collected gas. (Hint: Dalton's Law.)

b) (2.G, 8.A, 8.G, 9.A) Calculate the mass of potassium chlorate in the sample that decomposed.

5. (2.G, 8.A, 8.G, 9.A) Ammonium sulfate, an important fertilizer, can be prepared by the synthesis reaction of ammonia gas with aqueous sulfuric acid.

Write the balanced chemical equation here, including the state signifiers.

a) Calculate the volume of ammonia gas, in liters, needed at 20.℃ and 25.0 atm to react with 150 kg of sulfuric acid.

6. One method used in the eighteenth century to generate hydrogen was to pass steam through red-hot steel tubes. The following reaction takes place:

$3\ Fe_{(s)} + 4\ H_2O_{(g)} \rightarrow Fe_3O_{4\ (s)} + 4\ H_{2\ (g)}$

a) (2.G, 8.A, 8.G, 9.A) What volume of hydrogen at STP can be produced by the reaction of 6.28 g of iron?

b) 2.G, 8.A, 8.G) If the actual yield of hydrogen gas from the previous question is 3.00 L, calculate the percent yield of hydrogen gas.

c) (2.G, 8.A, 8.G, 9.A) What mass of iron will react with 500. L of steam at 250.°C and 1.00 atm pressure?

Challenge Questions:

7. (2.G, 8.A, 8.G, 9.A) Ammonium carbonate decomposes upon heating to produce ammonia gas, carbon dioxide gas, and water vapor. Calculate the *total* volume of gas produced at 22.0°C and 1.02 atm by the complete decomposition of 11.83 g of ammonium carbonate.

8. (2.G, 8.A, 8.G, 9.A) Pennies are composed of zinc coated with copper. A student determines the mass of a penny to be 2.482 g and then scratches the copper coating to expose the zinc. She puts the scratched penny in hydrochloric acid. A single replacement reaction occurs.

Write the balanced chemical equation here, including the state signifiers.

The student collects the hydrogen produced over water at 25.0°C. The collected gas occupies a volume of 0.899 L at a total pressure of 791 mmHg. The vapor pressure of water at this temperature is 23.78 mmHg. Calculate the percent zinc by mass in the penny.

Lesson 7.06 Student Notes:

7.07: Limiting and Excess Gas Stoichiometry

1. (3.B) Write a (*short*) argumentative paragraph about whether method a) or method b) is "better" for solving limiting and excess gas stoichiometry problems. (You decide what "better" means but explain that meaning.)

2. A 14.6 g sample of oxygen gas is placed in a sealed container with 2.5 g of hydrogen gas. The mixture is sparked, producing water vapor, in a synthesis reaction.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G, 8.H, 9.A) Calculate the volume of water vapor formed at STP.

b) (2.G, 8.G, 8.H, 9.A) Calculate the mass of excess reactant that remains.

3. Acrylonitrile, C_3H_3N, is the starting material for the production of acrylic fiber. It can be made from propylene, C_3H_6, by reaction with nitric oxide, NO, as follows:

$$2\ C_3H_{6\ (g)} + 2\ NO_{\ (g)} \rightarrow 2\ C_3H_3N_{\ (s)} + 2\ H_2O_{\ (l)} + H_{2\ (g)}$$

a) (2.G, 8.G, 8.H, 9.A) What is the theoretical yield of acrylonitrile when 21.6 g of propylene reacts with 21.6 g of nitric oxide?

b) (2.G, 8.G, 8.H, 9.A) Calculate the theoretical yield, in L, of hydrogen gas collected at 25.°C and 100. kPa?

c) (2.G, 8.G, 8.H, 9.A) What mass of excess reactant is left over?

4. One of the components of the fuel mixture on the Apollo lunar module involved a reaction with hydrazine, N_2H_4, and dinitrogen tetroxide, N_2O_4. If the balanced equation for this reaction is:

$$2\ N_2H_{4\ (l)} + N_2O_{4\ (g)} \rightarrow 3\ N_{2\ (g)} + 4\ H_2O_{\ (g)}$$

a) (2.G, 8.G, 8.H, 9.A) What volume of N_2 gas (measured at STP) would result from the reaction of 1500. Kg of hydrazine and 1000. kg of N_2O_4?

b) (2.G, 9.A) Be careful! What would be the volume of the N_2 gas after it is produced, then transferred to a sealed container at 2.00 atm and 20.0°C?

c) (2.G, 8.G, 8.H, 9.A) What mass of excess reactant is left over?

5. The synthesis reaction between sulfur dioxide gas and oxygen gas, which produces sulfur trioxide gas.

Write the balanced chemical equation here, including the state signifiers.

a) (2.G, 8.G, 8.H, 9.A) If 285.5 mL of sulfur dioxide reacts with 158.9 mL of oxygen (both measured at 315 K and 50.0 mmHg), what is the limiting reactant and the theoretical yield of the product?

b) (2.G, 8.G) If 187.2 mL of sulfur trioxide is collected (measured at 315 K and 50.0 mmHg), what is the percent yield for the reaction?

c) (2.G, 8.G, 8.H, 9.A) What volume of excess reactant is left over? (Assume it is measured at 315 K and 50.0 mmHg.)

Lesson 7.07 Student Notes:

7.08: Gas Stoichiometry Lab

Problem**:** Your goal is to evaluate the reaction of sodium bicarbonate (baking soda) with acetic acid (vinegar) as a potential replacement reaction for automobile airbags. Your task is to determine the correct ratio of baking soda and vinegar that leaves *no appreciable amount* of either reactant behind, yet fully inflates the bag (without bursting). The inflated bag is analogous to the airbags in cars.

Materials**:**

- sodium bicarbonate (baking soda) (10 g *available* per group)
- acetic acid (vinegar) (300 mL *available* per group)
- Ziploc™ gallon size bags
- graduated cylinders
- electronic balance
- weigh boats
- paper towels
- thermometer

Procedure:

Each group will design their own procedure. To keep it simple, determine the amount of sodium bicarbonate (baking soda) required to react with some small volume of vinegar (you choose the volume) so that the bag is inflated, with no appreciable quantity of either reactant left. To calculate the moles of vinegar present, multiply your chosen volume in mL by $\mathbf{8.33 \times 10^{-4}}$. Increase the sodium bicarbonate by 0.5 gram increments during your testing. You must have at least six trials completed.

You also need to determine the room temperature; since you are collecting a gas over water (and the vinegar does contain water), get the temperature and then look up the water vapor pressure at that temperature.

Some things to consider:

- How will you determine the volume of the bag before beginning the experiment?
- How will you get both reagents inside the bag without losing gas?

If you need other materials for your procedure, check with your teacher. Write your plan, showing all calculations before beginning the experiment. You may do some preliminary testing, but you must write them up as well.

Once you have determined the correct ratio of reactants that completely fill the bag without bursting, consult your teacher. Your data should include a well-designed plan, changes you made after the first trial, and results. Write up your plan below in a series of numbered, organized steps.

Our Plan**:**

Determining the volume of the bag:

Data: Pressure, Temperature, Volume

Atmospheric pressure today: (mmHg)	
Room temperature: (°C)	
Water vapor pressure at room temperature: (°C)	
Volume of empty zipper bag: (L)	

Reactants Data and Some Calculations:

To calculate the moles of vinegar present, multiply your volume in mL by **8.33×10^{-4}**.
Note in your test results whether the bag was comfortably inflated and whether all reactants appeared to be used up. You must do at least three tests; six lines are provided here. If your teacher wishes, you may use pH paper to test the liquid after each trial is done.

Trial	Vinegar		Baking Soda		Test Results
	Volume	# Moles	Mass	# Moles	
1					
2					
3					
4					
5					
6					

Analysis**:**

a) For your *best* trial, calculate the theoretical volume of carbon dioxide gas generated in the bag, in liters. Indicate which reactant is the limiting reactant. (Hint: Write a balanced equation first. Baking soda and vinegar react to form sodium acetate solution, carbon dioxide, and liquid water. You are doing a limiting/excess gas stoichiometry problem that also involves Dalton's Law.)

b) What is the mass of excess reactant left over for the best trial? (To calculate the mass of the vinegar, assume that the density of vinegar is the same as that of water.)

c) Compare your excess reactant (that you calculated) versus your test results. Do your test results support your identification of the excess reactant? Explain your answer.

d) Comment on the practicality of using a baking soda and vinegar system to generate air for car airbags. Consider cost, storage, and design in your (thoughtful, well written) analysis.

e) Use your partial pressure of $CO_{2\,(g)}$, volume, and temperature to verify that the volume of 1 mole of $CO_{2\,(g)}$ at STP is 22.4 L. (Hint: Combined Gas Law)

Lesson 7.08 Student Notes:

Unit 8: Solutions, Acids and Bases

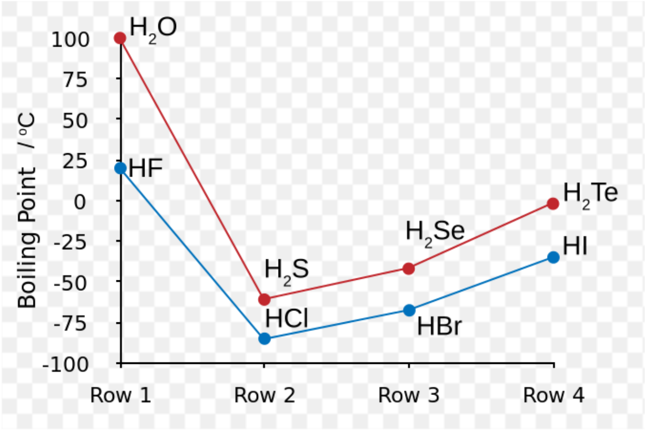
8.01: Intermolecular Forces / Don't Flip Your Lid Lab

1. (2.H, 2.I, 3.B) Examine the chart below comparing the boiling points of the noble gases. Explain the boiling point trend you observe here in terms of atomic structure, size, and polarizability, in a well-written paragraph.

Noble Gas	Boiling Point (K)
He	4.2
Ne	27
Ar	87
Kr	120
Xe	165

2. (2.H, 2.I, 3.B) $HCl_{(g)}$ has a solubility in water of 720 g/L at 20°C. This is much higher than the solubility of $Cl_{2\ (g)}$ in water, which is 7.5 g/L at 20°C. Explain the trend you observe here, using scientific principles, in a well-written paragraph. You must refer to both substances in your answer.

3. (2.H, 2.I, 3.B) Explain the abnormally high boiling points of $H_2O_{(l)}$ and $HF_{(l)}$, as depicted in the graph below, in terms of scientific concepts, in a well-written paragraph. The "row" stands for periods in the periodic table (i.e., Row "1" is the second period).



[14]

4. (2.H, 2.I, 3.B) Indicate which member of each pair would have the higher melting point based on their dominant IMFs. Explain your reasoning, making references to the IMFs in *both* substances in your answer and using scientific principles. You may have to draw Lewis structures.

a) $N_{2\,(s)}$ vs $NF_{3\,(s)}$

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[14] "File:Boiling-points Chalcogen-Halogen.svg." https://commons.wikimedia.org/wiki/File:Boiling-points_Chalcogen-Halogen.svg. Wikimedia Commons, 8 Feb 2012. Licensed under CC-BY-SA 3.0: https://creativecommons.org/licenses/by-sa/3.0/legalcode.

b) $NF_{3\ (s)}$ vs $NH_{3\ (s)}$

c) $C_2H_{6\ (s)}$ vs $C_3H_{8\ (s)}$

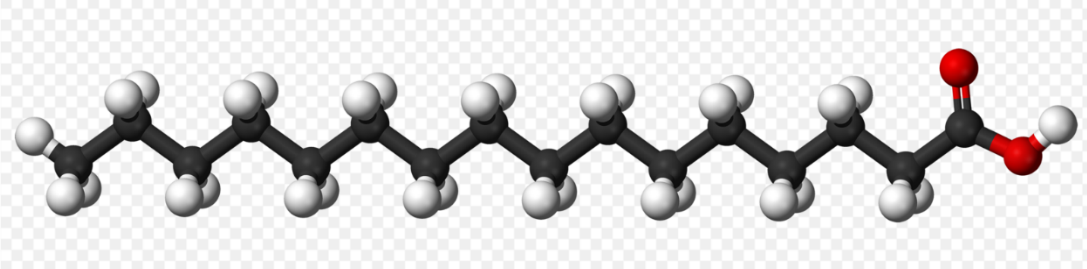
Challenge Problem:

5. (3.B, 3.F) Adhesives, such as glue and tape, rely on the power of intermolecular forces. So do many organisms. Research "gecko tape," explain how geckos can adhere to walls and surfaces using your knowledge of intermolecular forces and describe how scientists are building on this knowledge to develop new materials for practical use.

6. (3.B) Read the following passage and answer the question that follows.

Organic molecules vary in their solubility. For organic molecules, you need to examine the molecule's structure and see which part(s) are the largest. Fatty acids, for example, are mostly nonpolar because a large percentage of its molecules are nonpolar, consisting of many symmetrical carbon-to-hydrogen bonds. This fatty acid, palmitic acid, does contain polar parts that can form hydrogen bonds, but because there are so many carbon-hydrogen bonds those areas "overrule" the polar parts.

In general, the strongest polar bonds in organic molecules are (in descending order of strength) O-H, N-H, C-O and C-N. An organic molecule with many O-H bonds will dissolve very well in water for this reason. The molecule below, a fatty acid, [15] is mostly nonpolar because of its many C-H bonds "overruling" the tiny polar part on the right.



Indicate which member of each pair would have the higher melting point based on their dominant IMFs. Explain your reasoning, making references to the IMFs in *both* substances in your answer and using scientific principles.

a) $C_3H_7OH_{(g)}$ vs. $C_6H_{13}OH_{(g)}$

b) $CH_3OH_{(g)}$ vs. $CH_2(OH)_{2\,(g)}$

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[15] Mills, Jynto, and Mills, Ben. "File:Myristic-Acid-3D-Balls.png." *File:Myristic-Acid-3D-Balls.png*, Wikimedia Commons. 1 Aug. 2010, commons.wikimedia.org/wiki/File:Myristic-acid-3D-balls.png. Public domain.

Don’t Flip Your Lid Lab

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

Problem:

In this lab activity you will determine the relative melting points of four substances, $C_{25}H_{52}$ (paraffin), NaCl (table salt), I_2 (iodine) and $C_{12}H_{22}O_{11}$ (sucrose). Paraffin is a medium size nonpolar molecule. NaCl is an ionic compound. Iodine is a large nonpolar molecule, and $C_{12}H_{22}O_{11}$ is a large polar molecule.

Materials:

Hot plate – should be cold when you start

Solid samples of each chemical listed above

Aluminum foil

Scoopulas and tongs

Procedure:

1. In the space marked HYPOTHESIS, write a statement predicting in which order the salt, paraffin, sucrose, and glucose will melt.

2. Place the foil on the (cold) hot plate, and plug it in. Do not turn it on yet.

3. Transfer a small amount of each of the four samples into one of the four areas in a groove of the outer side of the foil. You only need a very small sample – just enough to be able to see the compound.

4. Turn the hot plate on to medium high and observe. You will need to record the relative order of melting. As soon as three of the compounds melt, turn off the hot plate.

5. Allow the foil to cool before using tongs to dispose of it in a trash can or as your teacher directs.

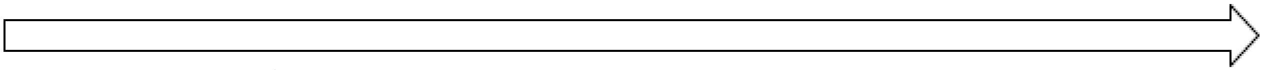
Hypothesis:

Observations*:* Record the relative order of melting for the four substances from lowest to highest melting point.

Analysis:

a) Was your hypothesis supported or unsupported by the experimental evidence? Explain your claim.

b) In the following table, write the names and formulas of the four substances from left to right in order of increasing strength of intermolecular forces. Then, complete the rest of the table.



Increasing strength of IMF

Name				
Formula				
Bond Type (ionic or covalent)				
Polarity (polar, nonpolar, or ionic)				
Total # Electrons				
Significant Intermolecular Force				

c) From what you know about attractive forces between molecules and/or ions, explain the relative order of melting points for EACH of the four substances. Consider bond types, relative strength, molecular polarity, electronegativity differences, polarizability, and the energy in J/mole required to melt the substance.

d) Hydrogen sulfide is a gas at room temperature, while water is a liquid, yet hydrogen sulfide has more electrons than water. Explain this anomaly.

e) In determining the strength of the intermolecular forces, it is important to consider the net forces present. Use this argument to defend why ice, despite having hydrogen bonding, has a much lower melting point than paraffin.

Lesson 8.01 Student Notes:

8.02: Solutions Concepts

1. (10.A) Describe the unique role of water in solutions in terms of polarity, in a well-written paragraph.

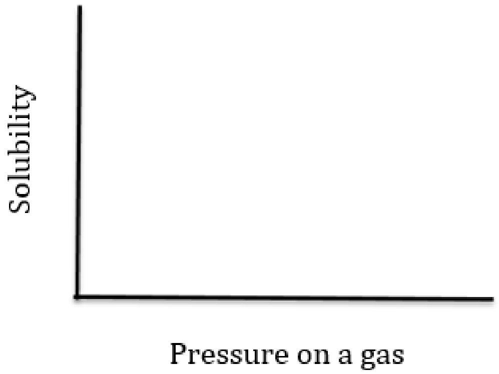
2. (3.B) Draw a particle diagram showing the solvation of one formula unit of $CaCl_2$ by water molecules, with the correct orientation. A formula unit of $CaCl_2$ has *three* ions. You must draw all three *separated* ions, each of them solvated by water molecules.

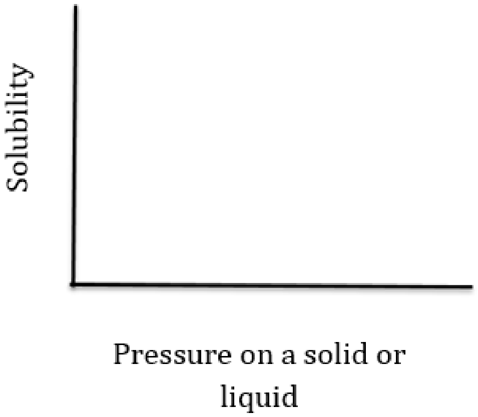
3. (3.B) Does each solute (the first of each pair) dissolve in the indicated solvent? Explain your answer in a well-written paragraph, using scientific principles. You must discuss both substances in your answers.

a) $CH_{4\ (g)}$ in $H_2O_{\ (l)}$

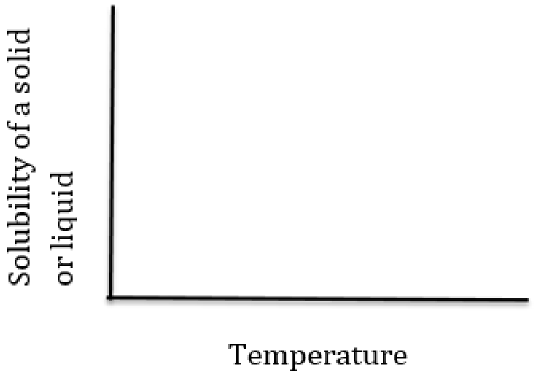
b) $HF_{\ (g)}$ in $H_2O_{\ (l)}$

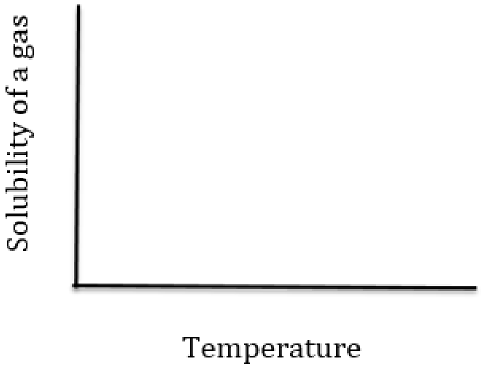
4. (3.B) On each set of labelled axes, draw a graph (just a line) that correctly represents each relationship.





5. (3.B) On each set of labelled axes, draw a graph (just a line) that correctly represents each relationship.





6. (3.B, 10.F) Cold “ice” water forms tiny bubbles as the water warms to room temperature. Explain this observation using the scientific principles discussed in this section.

7. (2.E, 3.B, 10.F) Design an experiment to dissolve 50. g of sugar cubes in 100 g of water as quickly as possible, using common laboratory equipment, in a step-by-step procedure.

8. (3.B, 10.E) Distinguish between these solutes based on electrical conductivity.

a) C_2H_5OH

b) H_2SO_4

c) H_2SO_3

d) $NaNO_3$

e) $FeCl_2$

f) $Ca(OH)_2$

g) $C_6H_{12}O_6$

h) RbOH

i) CO_2

Challenge Problems:

9. (3.A, 3.B, 3.C, 10.E) "Colloidal silver" is a common homeopathic remedy.

a) Let's say you had a sample of solution marked as "colloidal silver." Design a step-by-step procedure to help you determine whether this mixture is a colloid or not, using common laboratory equipment.

b) Research claims that state that colloidal silver is beneficial to health. Evaluate these claims, examining the data from both sides, for and against. In your opinion, is colloidal silver beneficial to human health?

10. (3.B, 3.F) Research the use of electrolytes in medicine and the food industry. Summarize your findings in a well-written paragraph.

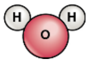
11. (2.H, 2.I, 3.B, 10.F) Because gases have almost no intermolecular forces, the solute-solvent interaction is very significant (more than normal) when trying to explain their solubility in various solvents.

a) The solubility of four gases *in water* measured at the same temperature and pressure decrease in this order: oxygen (O_2) > carbon monoxide > nitrogen (N_2) > helium (He). Explain this observation using your knowledge of intermolecular forces and solubility.

b) What would you predict is the order of solubility of oxygen, carbon monoxide, nitrogen, and helium in a nonpolar solvent, such as acetone or gasoline? Explain this observation using your knowledge of intermolecular forces and solubility.

Lesson 8.02 Student Notes:

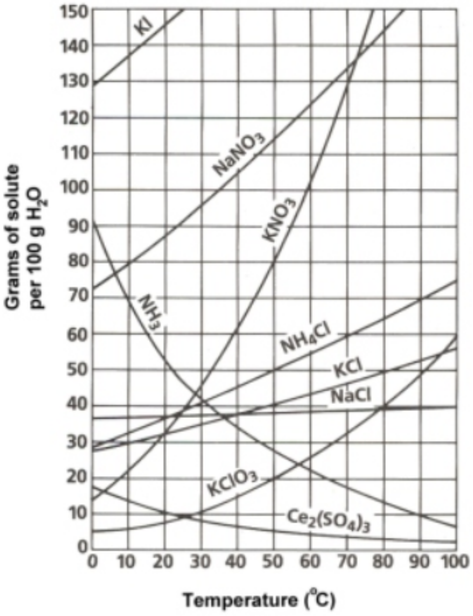
8.03: Saturation and Solubility Curves

1. (3.B, 10.E) Draw a labelled particle diagram depicting solute (assume NaCl) and solvent (assume water) in a beaker of saturated solution and in a beaker of unsaturated solution. For the Na^+ and Cl^- ions in both solutions, you must still show correct solvation by water molecules. Use individual labelled circles for each type of ion, and  for water.

Saturated solution

Unsaturated solution

2. [16](2.H, 2.I, 10.E) Based on the solubility curve, decide whether each of the following (1-4) is (A): unsaturated, (B): saturated, (C): supersaturated, or whether (D): not enough information is given.



a) 50 g KCl in 100 g of water at 90°C.

b) 50 g KCl in 100 g of water at 60°C.

c) 50 g KNO_3 in 25 g of water at 60°C.

d) 65 g KNO_3 in 400 g of water at 70°C.

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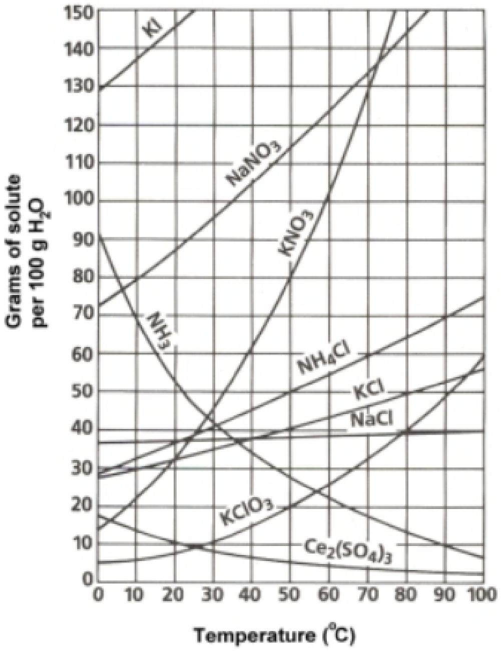
[16] "Solubility Curves." *3.1 Solubility Curves*, LibreTexts, 2019. Licensed under CC-BY-NC-SA 3.0 US: https://creativecommons.org/licenses/by-nc-sa/3.0/us/legalcode.

3. (2.H, 10.E) At which temperature will potassium nitrate and ammonia have the same solubility in water? What is that solubility? (Include the units.)



4. (2.H, 10.E) Which solid salt has the greatest increase in solubility from 20℃ to 70℃?

5. (2.H, 10.E) If a saturated ammonium chloride solution in 100 g of water at 30℃ is evaporated to dryness, how many grams will crystallize out?

6. (2.E, 2.H, 10.E) How many grams of potassium chlorate do you need to make a saturated solution in 100 g of water at 50°C?

a) Outline a procedure, in logical steps, to make this saturated potassium chlorate solution. Mention the correct use of the lab equipment you will need to make the solution.

7. [17](2.H, 10.E) You have a potassium chlorate solution containing 10 g of solute in 100 g of water at 50℃. What mass of solute do you have to add to make this solution saturated?

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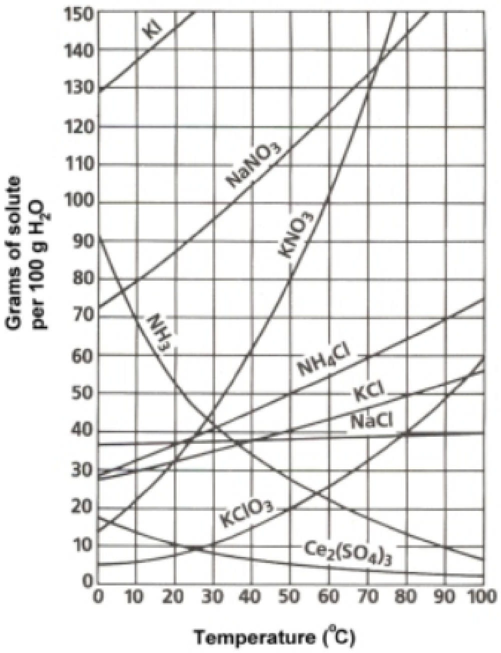
[17] "Solubility Curves." *3.1 Solubility Curves*, LibreTexts, 2019. Licensed under CC-BY-NC-SA 3.0 US: https://creativecommons.org/licenses/by-nc-sa/3.0/us/legalcode.



8. (2.E, 2.H, 10.E) You wish to make a supersaturated solution of 70 g ammonium chloride in 100 g of water at 80°C.

a) Outline a procedure, in logical steps, to make this supersaturated ammonium chloride solution. Mention the correct use of the lab equipment you will need to make the solution, and the temperature you will need to heat the water to.

b) [18]After the supersaturated ammonium chloride solution is made, how many grams of solute must crystallize out of the solution for it to become saturated at the same temperature, 80°C?

9. (2.H, 10.E) How much ammonia gas escapes out as a saturated solution in 100 g of water originally held at 40°C is warmed to 90°C?

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[18] "Solubility Curves." *3.1 Solubility Curves*, LibreTexts, 2019. Licensed under CC-BY-NC-SA 3.0 US: https://creativecommons.org/licenses/by-nc-sa/3.0/us/legalcode.

Challenge Question:

10. (3.B, 3.D) Supersaturated solutions are very important in the food industry. Research the uses and applications of supersaturated solutions in this field and write a summary that clearly explains the importance of the saturation concept to this industry.

Lesson 8.03 Student Notes:

8.04: Colligative Properties

1. (3.B) Common rock salt, NaCl (s), is the salt of choice for salting wintery roads.

a) Explain, in your own words, and using appropriate vocabulary related to this lesson, why we don't use sugar ($C_{12}H_{22}O_{11}$) to "salt" the roads.

b) Assuming cost is not a factor, suggest a salt (ionic compound) that would be an even more effective substance than NaCl for "salting" roads in the wintertime. Explain your choice.

2. (3.B) Water boils at about 68°C near the top of Mount Everest, the highest mountain in the world. Explain this observation in terms of the scientific principles covered in this section.

3. (2.G, 3.B) Calculate the molality of each solution.

a) 0.100 moles of diethyl ether ($C4H_{10}O$) dissolved in 0.250 kg water.

b) 30.0 g of chloroform ($CHCl_3$) dissolved in 650. g of water.

4. (2.G, 3.B) What mass of sucrose ($C_{12}H_{22}O_{11}$) do you need to make a 1.5 *m* sucrose solution in 500. g of water?

5. (4.B) Identify each physical property as intensive or extensive.

a) Freezing point depression and boiling point elevation
b) K_f and K_b
c) Molality
d) Freezing point and boiling point

6. (2.G, 3.B) The active ingredient in antifreeze is ethylene glycol, $C_2H_6O_2$. An ethylene glycol solution contains 21.2 g of ethylene glycol in 85.4 mL of water. Determine the freezing point and boiling point of the solution.

7. (2.G, 3.B) An aqueous solution containing 35.9 g of an unknown compound in 150.0 g of water has a freezing point of – 1.3 °C. Calculate the molar mass of the unknown compound in g/mol.

8. (2.G, 3.B) What mass of ethylene glycol ($C_2H_6O_2$), in grams, must you add to 1.0 kg of water to produce a solution that boils at 105.0°C?

9. (2.G, 3.B) Which aqueous solution has the highest boiling point: 0.5 *m* $C_6H_{12}O_{6\,(aq)}$, 0.5 *m* $NaCl_{\,(aq)}$, or 0.5 *m* $MgCl_{2\,(aq)}$? Explain your answer in a well-written paragraph *without* doing any calculations.

10. (2.G, 3.B) What mass of salt (NaCl) should you add to 1.00 L of water in an ice cream maker to make a solution that freezes at – 10.0°C? (The density of water is 1.00 g/mL.)

Challenge Problem:

11. (2.G, 3.B) An aqueous solution contains 8.92 g of MBr (M is a placeholder metal) in 500.0 mL of solution and has a freezing point of – 0.279°C. Calculate the molar mass of MBr, then use a periodic table to identify the element, M.

Lesson 8.04 Student Notes:

8.05: Molarity and Dilution; Solution Stoichiometry

1. (3.B) Draw a Venn Diagram, or another graphic organizer of your choosing, to show the similarities and differences between the concepts and formulas for molarity and molality. (See 8.04 if you need a molality refresher.)

2. (3.B) Describe at least two ways to make a solution *less* concentrated (that is, more dilute).

3. (2.G, 3.B) In order to maintain a sodium chloride concentration similar to ocean water, an aquarium must contain 3.6 g of sodium chloride per 100.0 g of water. What is the percent by mass of sodium chloride in the solution?

4. (2.G, 3.B) You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite is 3.62%. How many grams sodium hypochlorite are in the solution? How many grams of solvent are in the solution?

5. (2.G, 3.B, 10.C) Ascorbic acid, or vitamin C, is a water-soluble vitamin with a formula of $C_6H_8O_6$. A solution containing 80.5 g of ascorbic acid dissolved in 210 g of water has a density of 1.22 g/mL at 55°C. Calculate the molarity of ascorbic acid in this solution.

6. (2.G, 3.B, 10.C) The density of toluene, C_7H_8, is 0.867 g/mL, and the density of thiophene, C_4H_4S, is 1.065 g/mL. A solution is made by dissolving 8.10 g of thiophene in 250.0 mL of toluene. Calculate the molarity of thiophene in this solution.

(2.G, 3.B, 10.C) Calculate the molarity of each solution for the next four problems.

7. 3.0 mol sugar dissolved in 2.0 L of solution.

8. 0.030 moles potassium nitrate dissolved in 50.0 mL of solution.

9. 6.45 g of sodium sulfate dissolved in 250 mL of solution.

10. a) 0.250 kg of iron (III) nitrate dissolved in 1.50 L of solution

b) What is the molarity of iron (III) ion and the molarity of nitrate ion in the solution from a)?

(2.G, 3.B) Calculate the number of moles, mass of solute, or volume for each solution.

11. How many moles of ammonium phosphate are needed to make 150 mL of 3.0 *M* solution?

12. How many grams of sodium nitrate are needed to make 3.5 L of a 0.50 *M* solution?

13. How many grams of potassium chromate are needed to make 300.0 mL of a 1.25 *M* solution?

14. How many mL of 2.50 *M* tin (II) sulfate solution can be made using 1.8 g of this solute?

15. (2.E, 3.B) Sketch a cartoon that depicts the steps in preparing a standard solution. It does not have to be elaborate, but it must be accurate! (We will have the equipment available in class, so you can see what they look like.)

16. (2.G, 3.B, 10.D) Determine the concentrations for each of the following mixtures.

a) equal volumes of 3.0 *M* KCl & water (note: water is 0 *M* with respect to KCl):

b) equal volumes of 3.0 *M* KCl & 7.0 *M* KCl:

17. (2.G, 3.B, 10.D) Determine the concentrations of the following mixtures. All solutions are diluted with water.

a) 45 L of 3.6 *M* KCl diluted to a total volume of 60. L:

b) 215 mL of 2.8 *M* KCl diluted to a total volume of 270. mL:

c) 250. mL of 1.7 *M* NaCl diluted to a total volume of 300. mL:

d) 2.17 L of 0.81 *M* NaCl diluted to a total volume of 2.56 L:

18. (2.G, 3.B, 10.D) Determine the concentrations of the following mixtures.

a) 45 L of 3.6 *M* KCl and 71 L of *water*:

b) 215 mL of 2.8 *M* KCl and 47 mL *water*:

19. (2.G, 3.B) How much water do you have to add to each of these standard solutions to form a 0.50 *M* (diluted) solution?

a) 700. mL of a 1.25 *M* solution of potassium nitrate

b) 300. mL of a 0.85 *M* solution of copper (II) sulfate

20. (2.E, 3.B) Sketch a cartoon that depicts the steps in preparing a diluted solution. It does not have to be elaborate, but it must be accurate! (We will have the equipment available in class, so you can see what they look like.)

21. If 250. mL of 0.300 *M* sodium iodide reacts with excess lead (II) nitrate solution, how many grams of precipitate will form?

22. If 250. mL of 0.300 *M* sodium iodide reacts with 100. mL of 0.100 *M* lead (II) nitrate solution, how many grams of precipitate will form?

23. You mix 350. mL of 0.175 *M* lithium carbonate with 350. mL of 0.250 *M* barium nitrate. Calculate the theoretical yield of precipitate and the molarity of the excess reactant left over.

24. You mix 400. mL of 0.200 *M* iron (II) acetate with 250. mL of 0.300 *M* potassium phosphate.

a) What is the theoretical yield of the precipitate?

b) Outline a procedure to separate the precipitate from the supernatant in logical steps. Mention the following equipment and material: filter paper, funnel, beaker, stirring rod, distilled water.

c) The following is a data table taken from one trial of this experiment.

Mass of dry filter paper (g)	1.153 g
Mass of wet filter paper plus precipitate (g)	11.542 g
Mass of dry filter paper plus precipitate (g)	10.228 g

Calculate the actual yield of this trial. Then, calculate the percent yield of this trial.

d) What is the molarity of the excess reactant after precipitation has occurred?

Challenge Problems:

25. (2.G, 3.B, 10.C) Commercial concentrated aqueous ammonia is 28% NH_3 by mass and has a density of 0.90 g/mL. What is the molarity of this solution?

26. (2.G, 3.B, 10.D) Determine the final molarity of the mixture.

a) 45 mL of 3.6 *M* KCl and 71 mL of 6.2 *M* KCl:

b) 25 mL of 0.40 *M* NaCl and 53 mL of 0.25 M NaCl

27. (2.G, 3.B, 10.D) What volume of 2.5 *M* KCl must be added to 37 mL of 6.0 *M* KCl to make a final concentration of:

a) 3.5 *M*

b) 4.2 *M*

If your teacher so wishes, they may have you complete the following lab.

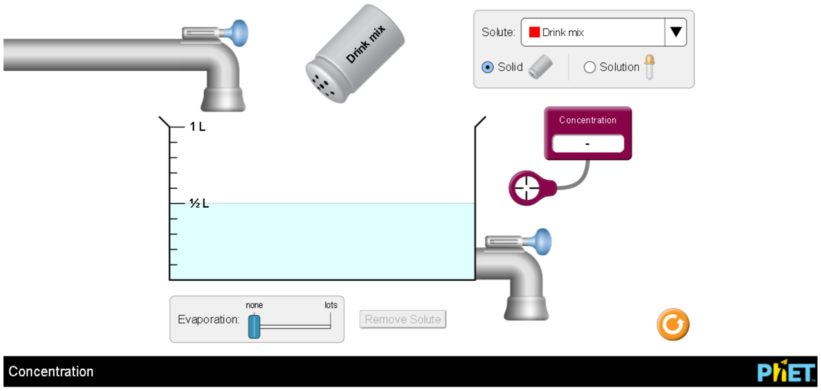
Molarity and Dilution Demonstration Lab

This lab uses the **Concentration** and **Molarity** simulations from PhET Interactive Simulations at University of Colorado Boulder, under the CC-BY 4.0 license.
https://phet.colorado.edu/sims/html/concentration/latest/concentration_en.html
https://phet.colorado.edu/sims/html/molarity/latest/molarity_en.html

Problem**:** It is possible to complete all sorts of molarity and dilution calculations and never fully understand these ideas conceptually. This online lab simulation will help you understand these concepts, describe the relationships between the variables that affect molarity, predict how solution concentration will change for any action (or combination of actions) that adds or removes water, solute, or solution, and explain why.

Procedure**:**

1. Explore the **Concentration** simulation. Find all the ways you can change the concentration of the solution in the beaker. For questions a)-d), leave the solute on “solid,” and move the target-shaped probe so it is “on top” of the liquid so you can measure the molarity.



Solute: Drink mix

● Solid ○ Solution

Analysis:

a) How does increasing and decreasing the mass of solute affect the molarity of the solution? Explain your answer, making specific references to what you did in the simulation, as well as the variables in the molarity formula.

b) Changing the solute (i.e., drink mix to cobalt (II) chloride) does *not* affect the molarity of the solution. Explain your answer, making specific references to what you did in the simulation, as well as the variables in the molarity formula.

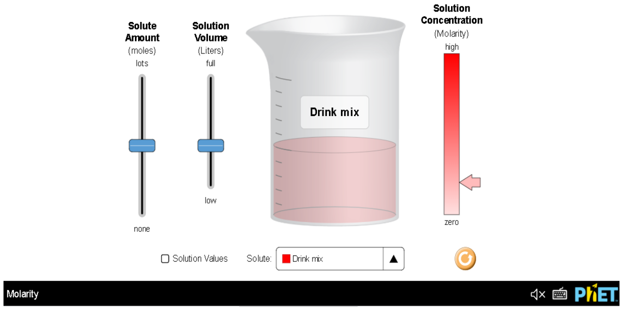
c) In the simulation, what are two different ways to change the molarity of the solution based on a volume change? How does each method increase or decrease the molarity? Explain your answer based on your knowledge of the molarity and/or dilution formulas.

d) How does pouring out a finished solution affect the molarity of the solution? Explain your answer, making specific references to what you did in the simulation, as well as the variables in the molarity formula.

e) Reset the simulation by clicking the orange button in the bottom right-hand corner. Move the target-shaped probe so it sits "on top" of the liquid. Set the solute to "solution." Add the "solution" solute to the liquid.

Describe what happens to the volume of the solution and the molarity of the solution as you add more "solution" solute. Explain your observation, making specific references to what you did in the simulation, as well as the variables in the molarity formula.

f) Explore the **Molarity** simulation and consider how it compares (i.e. similarities and differences) to the **Concentration simulation**. Write down your observations.



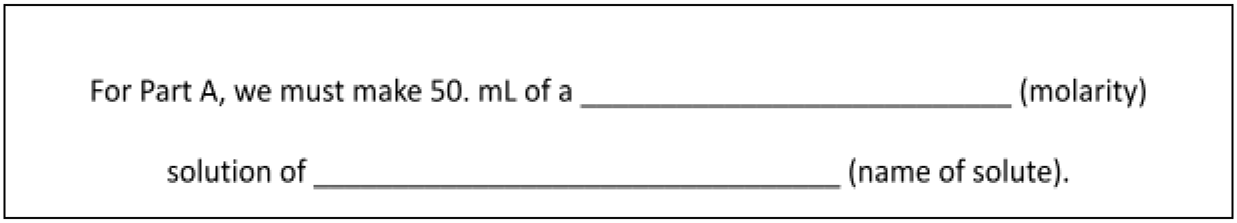
Lesson 8.05 Student Notes:

8.06: Preparation and Dilution of a Solution Lab

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

Part A Problem: Calculate the mass of solid solute needed to make 50. mL of solution of a specified molarity. Prepare this solution using the equipment that is provided at your station.

Before You Start: Get this information from your teacher and fill in the information below.



For Part A, we must make 50. mL of a ___________________________ (molarity)

solution of _________________________________ (name of solute).

Pre-Lab: If you are given a bottle of solid: examine the bottle to find its "F.W". "F.W" stands for "formula weight" and is the molar mass of the chemical in the bottle. Alternatively, you may be given the formula of the solute.

Either way, use the formula weight or the molar mass, plus what you have learned in the last two lessons, to calculate the mass of solid solute you need to make the solution described in the box. Show your calculations here.

Procedure: Using what you previously learned, write a step-by-step procedure for your group to follow in preparing this solution. Then, implement the procedure to make your Stock solution.

Part A Analysis:

How does each of the following sources of error impact the apparent molarity of your solution - does it make the molarity too high, too low, or is there no effect? Refer to the molarity formula in each of your answers in a well-written paragraph.

a) Spilling the weighed solid before you place it in the volumetric flask.

b) Pouring in too much water so that it goes past the mark on the neck of the flask.

c) Spilling the finished solution after it has been correctly made.

Show your solution and calculations to your teacher. Your teacher may test it to ensure it has been correctly made. Have them sign off here: ___________.

Part B Problem: Use your prepared Stock solution and the equipment at your table to produce 50. mL of a diluted solution of a specified molarity.

Before You Start: Get the relevant information from your teacher and fill in the information below for the first two blanks. You should already know what to put in the third blank.

For Part B, we must make 50. mL of a ___________________________ (molarity)

solution of _________________________________ (name of solute) using our prepared

Stock solution, which has a molarity of ________________________ (molarity).

Pre-Lab: Calculate the volume of prepared Stock solution you need to produce 50. mL of your assigned diluted molarity.

Procedure: Using what you previously learned, write a step-by-step procedure for your group to follow in preparing this diluted solution. Then, implement the procedure to make the diluted solution.

*Before you begin making the diluted solution, pour your Stock solution into a labelled beaker. You need the empty volumetric flask to make the diluted solution.

Show your solution and calculations to your teacher. Your teacher may test the solution to ensure it is correctly made. Have them sign here: ___________________

Lesson 8.06 Student Notes:

8.07: Beer's Law and Spectrophotometry / Beer's Law Lab

1. (2.G, 3.B, 10.C) The molar absorptivity constant of a particular chemical is 1.5 L/mol·cm. What is the concentration of a solution made from this chemical that has an absorbance of 0.72 with a cell path length of 1.1 cm?

2. (2.G, 3.B, 10.C) The absorptivity constant of a particular chemical is 1.5 L/mol·cm. What is the concentration of a solution made from this chemical if a sample with 2.0 cm path length has an absorbance of 1.20?

3. (2.G, 3.B, 10.C) Determine the concentration of a tryptophan solution in a 0.5 cm cuvette if its percent transmittance is 28.8% and its molar absorptivity constant is 6.4×10^3 L/mol·cm. (Hint: convert % transmittance to a decimal, then calculate its absorbance.)

4. (2.H, 2.I, 3.B, 10.C) Use the calibration curve equation to solve the following problems.

a) What is the molarity of a solution if its absorbance is 0.3?

b) What is the molarity of a solution if its absorbance is 0.55?

c) What is the absorbance of a solution that has a molarity of 0.80 *M*?

5. (2.H, 2.I, 3.B, 10.C) LASA students collected water samples from their homes to perform spectrophotometric analysis. They created a set of standard solutions of known concentrations of lead (II) sulfate.

Concentration (mol/L)	Absorbance
0.1006	0.179
0.200	0.200
0.300	0.300
0.400	0.400
Unknown	0.436

a) Use Excel or Google Sheets to graph and determine the equation for the line of best fit. Determine the concentration of the unknown.

b) The EPA action level guideline for lead in drinking water is at a concentration of 0.0450 mol/L. Should the students be drinking their water? Explain your conclusion in a well-written paragraph.

6. (2.H, 2.I, 3.B, 10.C) Students graphed the absorbance and concentration of several copper (II) sulfate pentahydrate solutions. Their calibration curve yielded a linear equation of $y = 2.2x - 0.00691$.

Use this equation to determine the concentration of a solution if its absorbance is:

a) 0.080	b) 0.142	c) 0.190

If your teacher so wishes, they may have you complete the following lab.

Beer's Law (Online) Demonstration Lab

This lab uses the **Beer's Law** simulation from PhET Interactive Simulations at University of Colorado Boulder, under the CC-BY 4.0 license.

https://phet.colorado.edu/sims/html/beers-law-lab/latest/beers-law-lab_en.html

Procedure:

1. Go to the Beer's Law simulation page.

2. Fix your settings: Leave the wavelength on "preset", click on the "absorbance" circle, and leave the solution to "drink mix". Move the green circular probe so it sits "on top" of the red liquid in the cuvette. Click the red button and move the probe again so it is blocking the green light from passing out of the other side. Leave the concentration alone for now. You should now see a numerical absorbance value in the green rectangle.

Analysis:

3. Move the double-headed orange arrow back and forth so it increases, then decreases, the path length.

a) How does an increase and a decrease in the path length affect the absorbance? Explain your answer using your knowledge of the absorbance formula.

4. Reset the simulation by clicking on the orange circle in the bottom right-hand corner. Fix your settings as described in step #2, above. Now move the concentration slider back and forth so it increases, then decreases, the concentration.

b) How does an increase and a decrease in the concentration affect the absorbance? Explain your answer using your knowledge of the absorbance formula.

5. Reset the simulation by clicking on the orange circle in the bottom right hand corner. Fix your settings as described in step #2, above. This time, you may choose the "solution", but leave the "wavelength" on "preset."

c) What solution did you choose? ____

6. Move the concentration slider to the right, gradually. At different positions – you can choose when - record the concentration and the resulting absorbance. Make sure your concentrations gradually increase as you read down the concentration column. (One point is done for you.)

d)

Concentration (mM)	Absorbance
0	0

7. In Excel or Google Sheets, graph your absorbance vs. concentration data.

e) What is the equation of your best fit line (the trendline)?

f) Use this best fit equation to calculate the concentration of the solution when the absorbance is 1.30. Show all the work.

8. Describe the effect of each source of error on the apparent concentration of your unknown. (Is the apparent concentration too high, too low, or is there no effect?) Explain each of your answers with specific reference to the concepts developed in this lesson.

g) You do not wipe off your fingerprints before placing a filled cuvette in the colorimeter.

h) You spill a cuvette, holding the unknown, on the table, before placing it in the colorimeter. (You still have enough solution for the light to pass through the sample.)

i) After you spill a cuvette (holding the unknown) all over the table, you top it off with distilled water before placing the cuvette in the colorimeter.

If your teacher so wishes, they may have you complete the following lab.

Beer's Law (In Person) Demonstration Lab

Problem: Obtain the absorbance of several standard solutions, graph absorbance vs concentration, then obtain the absorbance and concentration of an unknown.

Materials:

Vernier LabPro and colorimeter
Cuvettes
Distilled water
Standard solutions
Unknown solution

Before starting, ensure you are demonstrating laboratory safe practice, know the hazards of the chemicals involved. Read the procedure thoroughly and/or discuss the use and conservation of our lab resources and the correct disposal procedures.

Procedure: You may do this in a small group under your teacher's supervision.

1. Plug in and turn on the Vernier LabPro and colorimeter (set to Channel 1).

2. Create a blank by adding a small amount of distilled water to a cuvette. Place a lid on the cuvette. Wipe the smooth sides down with paper towel. Set this cuvette aside.

3. On the Vernier Lab Pro, set the absorbance to the wavelength given to you by your teacher.

4. Holding the cuvette by its *ridged* sides, place the blank in the sample holder of the colorimeter. Slide the holder lid closed. Press CAL to set to zero. Then, remove the blank.

5. Fill other empty cuvettes with small samples of your standard solutions. *Keep track of which is which!* Record the absorbance of each standard along with their concentrations.

6. Fill another empty cuvette with a sample of your unknown solution. Obtain and record its absorbance in your data chart.

Data:

Concentration (*M*)	Absorbance
0	
0.20	
0.40	
0.60	
0.80	
Unknown	

Analysis:

a) In Excel or Google Sheets, graph your absorbance vs. concentration data. What is the equation of your best fit line (the trendline)?

b) Use this best fit equation to calculate the concentration of the unknown. Show all the work. What is the concentration of the unknown?

c) Describe the effect of each source of error on the apparent concentration of your unknown. (Is the apparent concentration too high, too low, or is there no effect?) Explain each of your answers with specific reference to the concepts developed in this lesson.

- You do not wipe off your fingerprints before placing a filled cuvette in the colorimeter.

- You spill a cuvette, holding the unknown, on the table, before placing it in the colorimeter. (You still have enough solution for the light to pass through the sample.)

- After you spill a cuvette (holding the unknown) all over the table, you top it off with distilled water before placing the cuvette in the colorimeter.

Lesson 8.07 Student Notes:

8.08: Properties and Theories of Acids and Bases

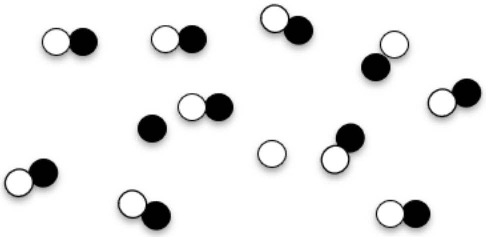
1. (10.E) Use the following legend to draw a particle diagram of each aqueous solution. For clarity, you do not have to draw the water molecules as part of the solution. You can fill up the box if you like.

○ = H^+ ○● = OH^-

An acid	A neutral solution	A base

2. (10.E) Draw a particle diagram depicting each aqueous solution. One has been done for you. You do not have to show hydration by the water molecules.

○ = H^+ ● = A^- ○● = HA ○

A concentrated strong acid	A dilute strong acid
	
A concentrated weak acid	A dilute weak acid

3. (7.B, 10.E) Examine the table below listing the strong acids, strong bases, and examples of weak acids. Write the correct chemical formula for each one. The weak bases have been completed for you.

Acids

Strong Acids

Hydroiodic acid	
Hydrobromic acid	
Hydrochloric acid	
Perchloric acid	
Chloric acid	
Nitric acid	
Sulfuric acid	

Bases

Strong Bases

All group 1/IA hydroxides and most group 2/IIA hydroxides such as…	
Calcium hydroxide	
Barium hydroxide	
Strontium hydroxide	

Weak Acids (examples)

Acetic acid	
Phosphoric acid	
Hydrofluoric acid	
Nitrous acid	

Weak Bases (examples)

Ammonia	NH_3
Methylamine	CH_3NH_2
Pyridine	C_5H_5N

4. (10.E) Examine the table on the previous page listing the strong acids, strong bases, and examples of weak acids again. Now indicate whether the acid or base is monoprotic (M), diprotic (D), or polyprotic (P) by adding the indicator next the name as follows in the table above: sulfuric acid (D) – meaning diprotic.

5. (8.E, 10.F) Write complete ionization equations (for the Arrhenius acids) or dissociation equations (for the Arrhenius bases). Don’t forget to put charges and states!

hypochlorous acid

6. (8.E, 10.F) Write complete ionization equations (for the Arrhenius acids) or dissociation equations (for the Arrhenius bases). Don't forget to put charges and states!

oxalic acid

7. (8.E, 10.F) Write complete ionization equations (for the Arrhenius acids) or dissociation equations (for the Arrhenius bases). Don't forget to put charges and states!

rubidium hydroxide

8. (8.E, 10.F) Write complete ionization equations (for the Arrhenius acids) or dissociation equations (for the Arrhenius bases). Don't forget to put charges and states!

strontium hydroxide

9. (10.F, 10.G) Given the list of compounds below, circle all species that would be considered amphoteric.

$H_2PO_4^-$

N_3^-

hydrogen carbonate ion

carbonate ion

ammonium ion

10. (10.G) For the following Brønsted-Lowry reaction, identify the acid, the base, and both conjugates. Also, identify the conjugate acid-base pairs.

$$HCl_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightarrow \quad H_3O^+_{(aq)} \quad + \quad Cl^-_{(aq)}$$

11. (10.G) For the following Brønsted-Lowry reaction, identify the acid, the base, and both conjugates. Also, identify the conjugate acid-base pairs.

$$NH_4^+{}_{(aq)} \quad + \quad CN^-_{(aq)} \quad \rightarrow \quad HCN_{(aq)} \quad + \quad NH_3{}_{(aq)}$$

12. (10.G) For the following Brønsted-Lowry reaction, identify the acid, the base, and both conjugates. Also, identify the conjugate acid-base pairs.

$$H_2SO_4{}_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightarrow \quad H_3O^+_{(aq)} \quad + \quad HSO_4^-{}_{(aq)}$$

13. (10.G) For the following Brønsted-Lowry reaction, identify the acid, the base, and both conjugates. Also, identify the conjugate acid-base pairs.

$$CO_3^{2-}{}_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightarrow \quad HCO_3^-{}_{(aq)} \quad + \quad OH^-_{(aq)}$$

14. (10.G) Are the following conjugate acid-base pairs, or not? If so, identify the acid and the base in the pair. All species are aqueous.

H_2O and H_3O^+ H_2O and O_2 SO_4^{2-} and HSO_4^- HCO_3^- and CO_2 NH_4^+ and NH_2^-

15. (8.E, 10.G) Using your knowledge of the Brønsted-Lowry theory of acids and bases, write equations for the following acid-base reactions and indicate each conjugate acid-base pair.

$CH_3NH_{2\,(aq)} + H_2O_{(l)} \rightarrow$

16. (8.E, 10.G) Using your knowledge of the Brønsted-Lowry theory of acids and bases, write equations for the following acid-base reactions and indicate each conjugate acid-base pair.

$OH^-_{(aq)} + HPO_4^{-2}{}_{(aq)} \rightarrow$

Lesson 8.08 Student Notes:

8.09: Equilibrium Concepts and Acid-Base Strength

1. (3.B) Write the equilibrium expressions for each of the following acid-base reactions.

$HCOOH_{(aq)} + H_2O_{(l)} \leftrightarrow HCOO^-_{(aq)} + H_3O^+_{(aq)}$

2. (3.B) Write the equilibrium expressions for each of the following acid-base reactions.

$HCO_3^-{}_{(aq)} + H_2O_{(l)} \leftrightarrow CO_3^-{}_{(aq)} + H_3O^+_{(aq)}$

3. (3.B) Write the equilibrium expressions for each of the following acid-base reactions.

$NH_{3\,(aq)} + H_2O_{(l)} \leftrightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$

4. (3.B, 10.E) Determine the stronger electrolyte in each pair of chemicals, based on their K_a or K_b values.

a) ammonium ion ($K_a = 5.6 \times 10^{-10}$) vs phosphoric acid ($K_a = 7.5 \times 10^{-3}$)

b) nitrite ion ($K_b = 2.2 \times 10^{-11}$) vs. cyanide ion ($K_b = 2.5 \times 10^{-5}$)

c) hydrogen sulfite ion ($K_a = 6.2 \times 10^{-8}$) vs formic acid ($K_a = 1.8 \times 10^{-4}$)

d) carbonate ion ($K_b = 2.1 \times 10^{-6}$) vs hydrogen phosphate ion ($K_b = 1.6 \times 10^{-7}$)

5. (3.B, 10.E) Arrange the acids in each group in order from lowest to highest K_a value. Explain each of your answers in detail and in terms of chemical principles (chemical bonds, bond strength, electronegativity, etc.).

a) HNO_2, HNO_3

b) $HClO_3$ (Cl = chlorine), $HBrO_3$, HIO_3

c) H_2SO_3, H_2SO_4

d) HOCl (Cl = chlorine), HOI, HOF, HOBr

e) HPO_4^-, $H_2PO_4^-$, H_3PO_4

f) OH^-, H_2O, H_3O^+

Acid-Base Equilibrium Simulation Lab

In the “Introduction” tab, take a few minutes to **play** with the simulation so you understand how it works. (http://phet.colorado.edu/en/simulation/acid-base-solutions)

6. (10.E) Fill out the following chart.

Systems	Equation	Graph (sketch it)	Results of using the tools
Water			
Strong Acid (HA)			
Weak Acid (HA)			
Strong Base (MOH)			
Weak Base (B)			

7. (2.H, 2.I) Explain why the graphs for strong acids and bases look fundamentally different from the graphs for weak acids and bases. Use scientific principles in your answer.

8. (3.B, 10.E) Write the K_c expression for each of the three systems (from the chart in #6) that demonstrate chemical equilibrium. Do not put H_2O (l) in your K_c expressions and use square brackets around each species to represent molarity.

   K_c = K_c = K_c =

9. (2.H, 2.I) *Explain* why the K_c values for each of the three systems in d) would be small (i.e., < 1) in terms of the ion concentrations, products, and reactants. Use detail!

Lesson 8.09 Student Notes:

8.10: Le Chatelier's Principle

1. (2.H, 2.I, 3.B) For the following reaction:

$5\ CO_{(g)} + I_2O_{5\ (s)}$ □ □ $I_{2\ (g)} + 5\ CO_{2\ (g)}$ $\Delta H = -\ 1175$ kJ

Predict the equilibrium shift and the effect on the indicated quantity for each change listed.

	Change	Direction of Shift (→; ←; or no change)	What is the effect on…	Effect (increase, decrease, no change)
A	Decrease in volume		Amount of CO_2	
B	Raise temp		Amount of CO	
C	Addition of CO		Amount of CO_2	
D	Removal of CO_2		Amount of I_2	

2. (2.H, 2.I, 3.B) In which direction will the equilibrium shift in response to each change, and what will be the effect on the indicated quantity?

$Ni_{(s)} + 4\ CO_{(g)} \leftrightarrow Ni(CO)_{4\ (g)}$ $\Delta H = -\ 161$ kJ

	Change	Direction of Shift (→; ←; or no change)	Effect on Quantity	Effect (increase, decrease, no change)
A	Add $Ni_{(s)}$		$Ni(CO)_4$	
B	Lower temp		CO	
C	Add CO		Amount of Ni	
D	Remove $Ni(CO)_4$		CO(g)	
E	Decrease in volume		$Ni(CO)_4$(g)	

3. (2.H, 2.I, 3.B) Use arrows (up/down) to indicate the effect of each of these disturbances (stresses) on the concentration of the reactants and products in this equilibrium:

$9\ kJ + 2\ SO_{2\,(g)} + O_{2\,(g)} \leftrightarrow 2\ SO_{3\,(g)}$

Disturbance	Effect on [SO_2]	Effect on [O_2]	Effect on [SO_3]	Direction of shift
Decrease [SO_2]				
Increase [O_2]				
Increase [SO_3]				
Increase pressure				
Decrease temp				

4. (2.H, 2.I, 3.B) Use arrows (up/down) to indicate the effect of each of these disturbances (stresses) on the concentration of the reactants and products in this equilibrium:

$N_{2\,(g)} + 3\ H_{2\,(g)} \leftrightarrow 2\ NH_{3\,(g)} + 100\ kJ$

disturbance	effect on [N_2]	effect on [H_2]	effect on [NH_3]	Direction of shift
Increase [N_2]				
Decrease [H_2]				
Increase [NH_3]				
Increase pressure				
Decrease temp				

Lesson 8.10 Student Notes:

Le Chatelier's Principle Demonstration

Objective: Predict and observe changes in equilibrium position when various stresses are placed upon a system at equilibrium. The system featured in this lab is:

$$\underset{\text{pink}}{Co\,(H_2O)_6^{2+}\,(aq)} + 4\,Cl^-\,(aq) \longleftrightarrow \underset{\text{blue}}{CoCl_4^{2-}\,(aq)} + 6\,H_2O\,(l)$$

Safety: Wear goggles! Follow your teacher's directions for disposal of all the chemicals.

Materials:

0.2 *M* cobalt (II) chloride	12 *M* HCl solution	0.1 *M* silver nitrate solution
Distilled water	400 mL beaker/hot water bath	400 mL beaker/ice bath
Test tube rack/4 test tubes/test tube tongs		Pipettes

Preparation and Procedure: *(You may complete this lab under your teacher's supervision.)*

Preparation

1. Label your test tubes P1, P2, B1 and B2.
2. To each of the four test tubes, add *about* 5 mL of cobalt (II) chloride solution.
3. Add about 5 mL of 12 *M* HCl solution to B1 and B2. These solutions should now be blue.
4. Set aside P1 and B1 as control solutions.
5. Set up a hot water bath (to about 80-85℃) and an ice water bath.

Procedure

6. Add about 5 mL of 12 *M* HCl solution to P2. Observe and record the changes.

7. Add about 5 mL of distilled water to P2. Observe and record the changes.

8. Add 0.1 *M* silver nitrate solution to B2. Observe and record the changes.

9. Get P2 and place it in the hot water bath. Wait, then observe and record the changes.

10. Use tongs to obtain P2 and put it in the ice water bath. Wait, observe, and record changes.

Data:

Change	Description	Color (before change)	Color (after change)
Procedure #6	P2 + HCl		
Procedure #7	P2 from #6 + H_2O		
Procedure #8	B2 + $AgNO_3$		
Procedure #9	P2 from #7 + heat		
Procedure #10	P2 from #9 + ice		

Analysis:

Explain each of the changes in Procedure #6-10 in terms of Le Chatelier’s Principle and chemical equilibrium.

a) Procedure #6

b) Procedure #7

c) Procedure #8

d) Procedure #9

e) Procedure #10

8.11: Strong Acids and Bases pH and pOH Calculations

1. (2.G, 10.C) Calculate the H^+ concentration given the following OH^-: Acidic/Basic/Neutral

a) $[OH^-] = 1 \times 10^{-13}$ *M*

b) $[OH^-] = 2.7 \times 10^{-4}$ *M*

c) $[OH^-] = 1 \times 10^{-3}$ *M*

d) $[OH^-] = 6.3 \times 10^{-10}$ *M*

2. (2.G, 10.C) Calculate the OH^- concentration given the following H^+: Acidic/Basic/Neutral

a) $[H^+] = 1 \times 10^{-3}$ *M*

b) $[H^+] = 3.6 \times 10^{-5}$ *M*

c) $[H^+] = 1 \times 10^{-2}$ *M*

d) $[H^+] = 7.8 \times 10^{-8}$ *M*

3. (2.G, 10.E) Go through each solution described in questions #1 and 2, and state if it is acidic, basic, or neutral.

4. (2.G, 10.H) A student performed the following calculations to determine the pH of solution with a $[OH^-] = 3.2 \times 10^{-4}$ M. Identify the error in the student's calculation and correct it to determine the correct pH.

$$[H^+] = (3.2 \times 10^{-4}) \div (1 \times 10^{-14}) = 3.2 \times 10^{10} \text{ M}$$

$$\text{pH} = -\log [3.2 \times 10^{10}] = -10.50$$

5. (2.G, 10.C, 10.H) Use your knowledge of pH to fill in the blanks in the following table. Complete ALL blanks. In the space below the table, show work for at least one blank answer from each column – pH, pOH, $[H^+]$ and $[OH^-]$.

pH	**pOH**	**$[H^+]$**	**$[OH^-]$**
		1×10^{-3}	
3.092			
			2.7×10^{-4}
8.319			
	6.942		
		3.6×10^{-5}	
			6.3×10^{-10}

6. (2.G, 8.E, 10.C, 10.H) What is the pH of a solution that contains 25 grams of hydrochloric acid dissolved in 1.5 liters of solution?

7. (2.G, 8.E, 10.C, 10.H) What is the pH and pOH of a solution that contains 1.32 g of sodium hydroxide dissolved in 750 mL of solution?

8. (2.G, 8.E, 10.C, 10.H) What is the pH of 1.0 L of 0.0100 *M* of hydroiodic acid? Assume complete ionization.

9. (2.G, 8.E, 10.C, 10.H) Calculate the pH of 500. mL of 0.0100 *M* of barium hydroxide.

10. (2.G, 8.E, 10.C, 10.H) What is the molarity of a perchloric acid solution with a pH of 1.45?

11. (2.G, 8.E, 10.C, 10.H) What is the molarity of a calcium hydroxide solution that has a pH of 10.07?

12. (2.G, 8.E, 10.C, 10.H) What is the pH of a solution that contains equal volumes of 0.050 *M* nitric acid and 0.100 *M* hydrochloric acid? (Hint: Review the molarity and dilution lesson…)

Challenge Question:

13. (2.G, 8.E, 10.C, 10.H) What is the pH of a solution that contains 25 mL of 0.200 *M* hydrochloric acid and 35 mL of 0.100 *M* sulfuric acid?

14. (2.G, 8.E, 10.C, 10.H) Up until this point you have been calculating the pH of strong acids or bases, but what about a weak acid or base? Watch this short video (https://bit.ly/2NEjmKl) and then calculate the pH of the following two solutions.

a) a 0.32 M solution of perchloric acid

b) a 0.32 M solution of propanoic acid ($K_a = 1.318 \times 10^{-5}$)

Lesson 8.11 Student Notes:

8.12: Neutralization, Titration and Titration Curves

1. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) Calculate the molarity of 75.0 mL hydrochloric acid that could neutralize 49.1 mL of 1.40 *M* $Ca(OH)_2$.

2. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) It was found that 33.8 mL of a triprotic acid was required to neutralize 43.1 mL of 1.10 *M* NaOH. What is the molarity of the acid? (Use "H_3A" for the acid formula.)

3. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) Industrially produced nitric acid is 8.0 *M*. What volume of nitric acid is needed to exactly neutralize 37.9 mL of a 3.50 *M* calcium hydroxide spill?

4. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) Perchloric acid is used in preparing rocket fuels. What volume of 5.0 *M* perchloric acid will completely react with 50.0 mL of a 1.50 M potassium hydroxide solution?

For each situation, calculate the molarity of the excess reactant and the pH of the resulting solution after neutralization has occurred.

5. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) 75.0 mL of 0.250 *M* hydrochloric acid and 50.0 mL of 0.150 *M* potassium hydroxide.

6. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) 65.0 mL of 0.300 *M* sulfuric acid and 80.0 mL of 0.800 *M* lithium hydroxide.

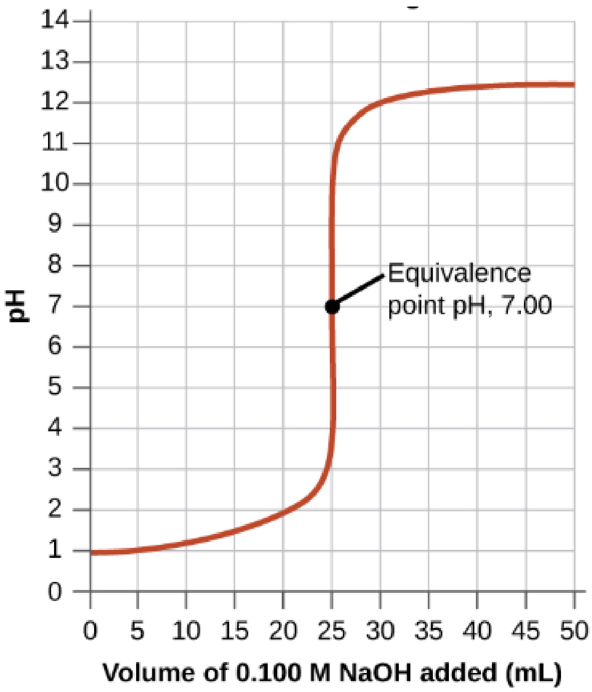
7. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) 0.50 *M* calcium hydroxide solution is used to titrate 75.20 mL of perchloric acid. What is the molarity of the acid?

	Volume (mL)
Initial volume	17.40 mL
Final volume	49.32 mL

8. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) The concentration of a solution of potassium hydroxide is determined by titration with sulfuric acid. A 25.0 mL sample of potassium hydroxide reaches endpoint when 35.6 mL of 0.562 *M* sulfuric acid is added. What is the concentration of the potassium hydroxide solution?

9. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) What is the concentration of acid in rainwater when 100.0 mL is titrated with 25.12 mL of 0.00105 *M* NaOH? Since acid rain contains several acids, assume the acid is monoprotic and use the symbol HA to represent the acids that are present.

10. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) For this titration curve[19], generated for the reaction of 20.00 mL of sulfuric acid with 0.100 *M* sodium hydroxide:



a) What is the volume of sodium hydroxide required to reach the equivalence point?

b) Calculate the molarity of the sulfuric acid in the flask.

11. (2.G, 7.B, 8.E, 8.G, 10.C, 10.D) 15.00 mL of calcium hydroxide in a flask is titrated with 0.0750 *M* nitric acid. As always, the pH at the equivalence point is 7. It takes 20.00 mL of nitric acid to neutralize the calcium hydroxide.

a) Label the axes, then sketch a titration curve that could describe this titration. Hint: You're *starting* with calcium hydroxide in the flask.

b) Calculate the molarity of the calcium hydroxide in the flask.

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[19] "Figure 1." *14.7 Acid-Base Titrations*, Rice University, opentextbc.ca/chemistry/chapter/14-7-acid-base-titrations/. Licensed under CC BY 4.0: https://creativecommons.org/licenses/by/4.0/legalcode.

Lesson 8.12 Student Notes:

8.13: Lab - Percent Acid in Vinegar / Percent Acid in Mustard

The following procedure is for the titration of vinegar with sodium hydroxide. If we have you titrate mustard, a separate procedure/assignment will be given to you to complete.

Problem*:* Titrate vinegar with sodium hydroxide and use the data to calculate the percent of acetic acid present in commercial vinegar.

Materials*:*

50 mL beakers
0.50 *M* sodium hydroxide
24-well microplate
White vinegar
Toothpick (for stirring)
Phenolphthalein
Plastic pipettes

Procedure:

1. Obtain about 10 mL of vinegar and about 20 mL of sodium hydroxide in two beakers. Label them to keep track of what solution is where.
2. Label two plastic pipettes “V” and “B” (for base).
3. Add 15 drops of vinegar to one well in your microplate using the “V” pipette.
4. Add 1 drop of phenolphthalein to the well containing the vinegar.
5. Begin adding sodium hydroxide, drop by drop, using the “B” pipette. Use the toothpick to stir well after every drop of base added. Keep adding drops of base until a *very faint* pink color persists in the well after 20 seconds, even after stirring.
6. Record the number of drops of base required to cause endpoint.
7. Complete two more trials using different wells.

Data (2.F): (Draw your own data table using a ruler.)

Analysis:

a) (2.G, 7.B, 8.E, 10.G) Write the balanced neutralization equation for the reaction of acetic acid with sodium hydroxide. All states must be present.

b) (2.G, 8.G, 10.C) For each trial, calculate the molarity of the acetic acid. You can use number of drops to stand in for volume.

c) (2.G) Use your results from above to calculate the average molarity of acetic acid in vinegar.

d) (2.G, 8.A, 8.C) Calculate the percent by mass of acetic acid in vinegar using the average molarity you calculate above. The density of the vinegar is 1.00 g/mL.

Hint: You will need to calculate the molar mass of pure acetic acid, then use molar mass to convert from # moles of acid (in the average molarity) to mass of acid. Also, refer to the molarity and percent by mass formulas.

e) (2.H, 2.I) Why can you use the number of drops to stand in for volume?

f) (2.H, 2.I) Why wouldn't a procedure like this work to determine the molarity of red wine vinegar?

8.14: Lab - Standardization of NaOH Titration

Problem: "Standardize" (determine the molarity of) a solution of sodium hydroxide solution by titration with standard 0.15 *M* hydrochloric acid. NOTE: Your teacher may revise this lab to include generating a titration curve.

Materials: (Per group)

Ring stand
Buret clamp
50 mL buret (check for leaks or cracks)
Erlenmeyer flask
Phenolphthalein indicator
Funnel (for sodium hydroxide/buret)
Standard 0.15 *M* hydrochloric acid
Unknown sodium hydroxide (A, B, or C)
Beakers to hold acid and base
Pipet filler
Graduated pipet
Deionized water in water bottle

Procedure:

1. You will be personally assigned an unknown concentration of **NaOH** (A, B, or C).

2. From the bottles, obtain some 0.15 *M* HCl in the red beaker and some of your assigned unknown NaOH in the green beaker. You will use the beaker solutions to refill your buret(s) while you do the lab.

3. If you are the first group to do this lab from your teacher's classes, rinse your buret with about 2 ~ 3 mL of sodium hydroxide. Pour the used rinse into the sink. **If there is solution already in the buret, do not rinse or empty it**.

4. CLOSE THE STOPCOCK ON THE BURET! (The handle will be horizontal when fully closed.)

5. Fill the buret with sodium hydroxide, using the funnel, and open the stopcock just long enough to let out air bubbles. **Pour slowly because it is very easy to overflow the burets and cause a chemical spill**.

You do NOT have to have the buret meniscus on 0.

6. Place the index card behind the buret. Carefully read the initial volume of sodium hydroxide in the buret and record it in the data table. RECORD ONE UNCERTAIN DIGIT (TWO DECIMAL PLACES), so all your volumes in this lab will be recorded to the hundredth place.

7. Use the pipet filler and pipet to get 10.0 mL of hydrochloric acid. Put this acid in the Erlenmeyer flask. Add two to three drops of phenolphthalein to the acid. The flask solution should be clear and colorless.

8. Raise the buret carefully and place the flask underneath the buret tip.

Time to titrate!

9. Put a white piece of paper underneath the Erlenmeyer flask (*not the index card!*)

10. Open the stopcock to allow approximately 6 mL of the base to flow into the flask. Have a second student continuously swirling the flask. Observe the color changes occurring.

11. Continue to add the NaOH slowly **while swirling the flask**.

The pink color will gradually last longer and longer. When you feel you are getting really close to the endpoint, turn the stopcock handle very carefully to deliver single drops, then fractions of a drop, to the flask.

You can knock off drops hanging off the buret tip by squirting distilled water to get the hanging drops to enter the flask or use the distilled water bottle to get NaOH drops to slide down the flask walls and into the liquid.

12. When a *very* pale “baby” pink color appears and *persists for 5 seconds or more* of swirling, you have reached the endpoint.

Close the stopcock. Record the final volume of the sodium hydroxide in the buret to ONE UNCERTAIN DIGIT.

Note: if you accidentally pass your endpoint (bright, dark magenta pink color appears and lasts), you can “**back titrate**” by pipetting more hydrochloric acid to the flask. This will cause the flask solution to become colorless again. Add this new volume to the old volume of acid and record the “new” final volume of acid with two decimal places. You can then add the NaOH slowly to the flask again and attempt once more to carefully reach the endpoint.

13. Pour the solution out of the flask and into the sink. Repeat the titration again from step #5 (if you need more buret solution). You should estimate how much sodium hydroxide you will need for this titration based on the information that you gathered from the first trial.

14. Complete one more trial, for a total of three, with the same unknown sodium hydroxide solution. If you have solution left over in either buret, you can leave them there - no need to empty them out.

Standardization of NaOH by Titration Lab Student Data

Assigned Unknown Letter: ______

Data (2.F):

	Trial 1	Trial 2	Trial 3
Volume of acid (mL)			
Initial buret reading (mL)			
Final buret reading (mL)			
Volume used (mL)			

Calculations (2.G, 8.E, 8.G, 10.C, 10.D): Show all work, and calculate the molarity of NaOH for each trial. **All work must be neat and easy to follow.** Record each answer to the appropriate number of sig figs and include the correct unit after every final answer.

Trial 1: Molarity of NaOH WITH UNIT:
Trial #2: Molarity of NaOH WITH UNIT:
Trial #3: Molarity of NaOH WITH UNIT:
Average molarity of NaOH WITH UNIT: